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Basic Western Saga

UNIT N
bases and acids

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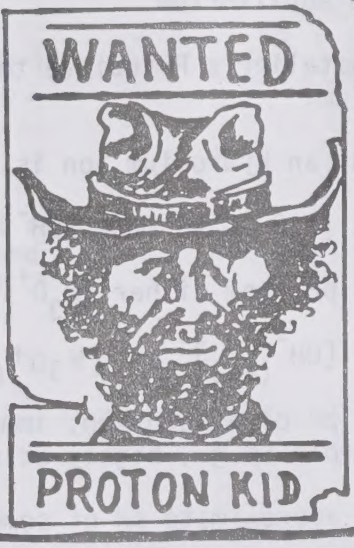
Wanted Proton Kid

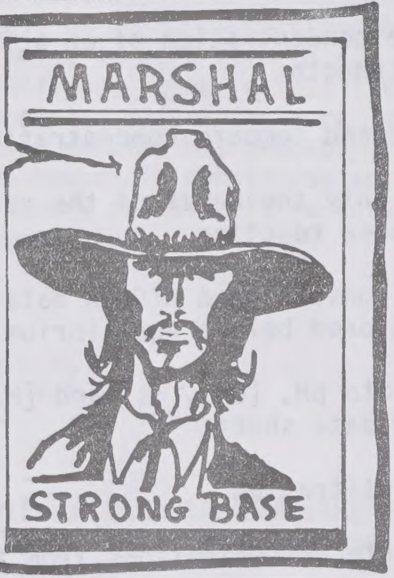
He has got everything. Everybody gets a charge out of him.



The Good Guy

MARSHAL STRONG BASE

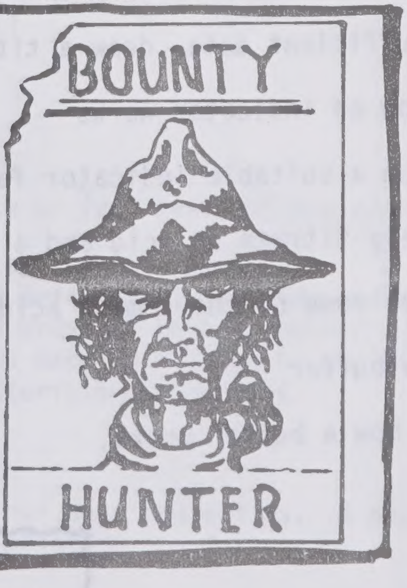
He is the only one who can get the Proton Kid. Why? He is a super guy, honest and courageous.



The Bad Guy

W. B. BOUNTY HUNTER

He aint got what it takes. Lack of character.



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ACIDS AND BASES

UNIT N

OBJECTIVES

Upon completion of Unit N the student should be able to:

N1. recognize acid and base properties

N2. identify any Bronsted-Lowry acid or base in a reaction equation

N3. define an equilibrium

N4. use Le Chatelier's Principle to predict the effect of an imposed change on an equilibrium.

N5. state what an hydronium ion is

N6. calculate $[H_3O^+_{(aq)}]$ from $[OH^-_{(aq)}]$ or $[OH^-_{(aq)}]$ from $[H_3O^+_{(aq)}]$

N7. calculate pH from either $[H_3O^+_{(aq)}]$ or $[OH^-_{(aq)}]$

N8. calculate $[OH^-_{(aq)}]$ or $[H_3O^+_{(aq)}]$ from pH

N9. given the pH of a solution, immediately classify the degree of acidity or basicity of the sample (e.g., highly acidic or basic, slightly acidic or basic, neutral,etc.)

N10. state the approximate pH of some common substances

N11. define concentration of an acid and strength of an acid and compare concentration and strength

N12. define and compare concentration of a base and strength of a base

N13. given only the names of the reactants, write a net ionic equation for any Bronsted-Lowry acid-base reaction

N14. after consulting a *ALCHEM* data sheet, predict whether the forward or reverse reaction was favored before equilibrium in a proton transfer reaction

N15. calculate pH, $[OH^-_{(aq)}]$ and $[H_3O^+_{(aq)}]$ of a 1.0 M solution of any weak acid on the *ALCHEM* data sheet

N16. define titration

N17. calculate concentrations from acid-base titration data obtained in the laboratory

N18. given sufficient data, draw a titration curve

N19. state how an indicator works

N20. decide on a suitable indicator for a given titration curve

N21. accurately titrate an acid and a base

N22. recognize some of the common acids and bases and know their uses

N23. define a buffer

N24. explain how a buffer works.

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ACIDS AND BASES

Chemical coupounds can usually be classified in one of three categories - ionic, molecular or acidic.

Ionic Compounds

Ionic compounds exist as solids at room temperature and form conducting solutions when dissolved in water (i.e., are electrolytes).

Example:

$$Mg(OH)_{2\,(s)} \longrightarrow Mg^{2+}_{(aq)} + 2\,OH^{-}_{(aq)}$$

exists as solid — dissociates into ions and therefore forms a conducting solution

Molecular Compounds

Molecular compounds can be solid, liquid or gas at room conditions, exist as molecules, and often do not dissolve in water. If they dissolve, undissociated molecules are usually found in solution (i.e., molecular compounds are nonelectrolytes).

Example:

$$C_{12}H_{22}O_{11(s)} \longrightarrow C_{12}H_{22}O_{11(aq)}$$

Acidic Compounds

Acidic compounds or acids possess properties of both previously mentioned categories. For example, many acids exist as molecules rather than ions, but form conducting solutions. All acidic compounds are acids in solution but only some exist in pure form (e.g., solid salicylic acid, liquid acetic acid and gaseous hydrogen sulfide). Other acids, like carbonic and sulfurous acids only exist in aqueous solution and decompose before the pure compound can be isolated. Chemically speaking, acids are interesting species but probably the best reason for studying acids is that a large number of common household substances are acidic. For example, vinegar contains ethanoic (acetic) acid and lemons contain citric acid.

All aqueous solutions of acids have the following properties:

a. conduct an electric current (are electrolytes)
b. turn blue litmus to red
*c. taste sour
d. react with active metals (e.g., Mg or Zn) to form H_2 gas
e. neutralize bases (make a base solution neutral - see below)

Basic Substances

Many common household substances are bases; e.g., milk of magnesia contains the basic substance magnesium hydroxide and household ammonia is a common basic cleaning agent.

All aqueous solutions of bases have the following properties:

a. conduct an electric current (are electrolytes)
b. turn red litmus to blue
*c. taste bitter
*d. feel slippery
e. neutralize acids

* As a general rule, chemists *never* attempt taste or feel tests unless *certain* about the nature of the substance.

The above lists of properties provide *operational* definitions of acids and bases by listing a number of operations than can be performed in order to decide whether a substance is an acid or a base. Note that operational definitions make no attempt to explain how the substances react or why the substances have their charateristic properties.

Neutral Solutions

A *neutral solution* is one that has neither acid nor base properties. A neutral solution can be formed by reacting a base solution with an acid solution.

ACIDS AND BASES

CLASSIFICATION OF SOLUTIONS - DEMO N1

Purpose:

To classify a number of solutions as acidic, basic, neutral ionic or neutral molecular.

Predemo Information:

Acid and base solutions have been defined earlier in this unit. A neutral ionic solution is a solution that is neither acid nor base but is still an electrolyte. A neutral molecular solution is neither acid nor base and not an electrolyte (i.e., it is formed by dissolving a molecular substance).

Materials:

1 - conductivity apparatus (light bulb or meter connected to electrode probes)

1 - tube blue litmus paper

1 - tube red litmus paper

1 - wash bottle containing distilled $H_2O_{(\ell)}$

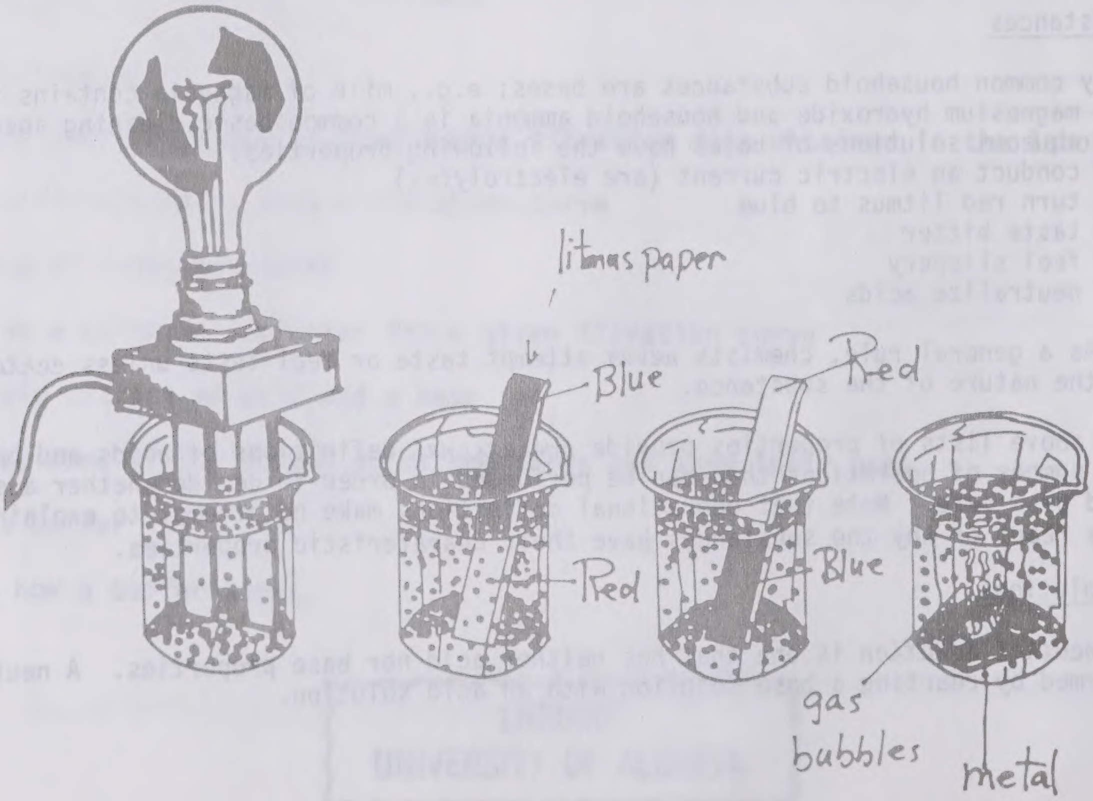
1 - 400 ml beaker (for waste)

10 - unknown solution samples in 50 ml beakers

10 - small pieces of $Mg_{(s)}$

Procedure: (Students record all observations.)

1. Number the unknown samples.
2. Test Solution 1 for conductivity. Rinse the electrodes with distilled H_2O after each test.
3. Test Solution 1 with a strip of blue litmus and a strip of red litmus.
4. Place a small piece of Mg into Solution 1 to test for production of H_2 gas.
5. Classify the solution on the basis of observations from these tests.
6. Repeat Steps 2 through 6 for each of the other unknown solutions.
7. Record the formula for each substance as given by the teacher.



ACIDS AND BASES

CLASSIFICATION OF SOLUTIONS - DEMO N1

Observations:

	Electrolyte (yes or no)	Blue Litmus (change or no change)	Red Litmus (change or no change)	Active Metal Test (gas or no gas)	Type of Solution (acid, base, neutral ionic, or neutral molecular)	Formula for Solute Tested (from teacher)
1						
2						
3						
4						
5						
6						
7						
8						
9						
10						
11						

Questions:

1. Which substances formed acidic solutions?

2. Are there any similarities among the compounds in Question 1 that would allow you to recognize an acid?

3. Which substances formed basic solutions?

4. Is there any way of recognizing a base?

CONCEPTUAL DEFINITION, BRONSTED-LOWRY THEORY

Operational definitions, such as those previously given for acids and bases, are obtained strictly from observation. An operational definition is both useful and accurate, but does not explain *why* substances have characteristic properties. Theories developed to explain why substances behave in a certain way are called *conceptual* definitions.

As the study of chemistry has progressed, several conceptual definitions have been developed for acids and bases. One of the most useful conceptual definitions, and the one that will be primarily used in *ALCHEM 30*, is the Bronsted-Lowry acid-base theory. The theory is named for the two scientists, J.N. Bronsted and T.M. Lowry, who developed their ideas independently in the same year - 1923.

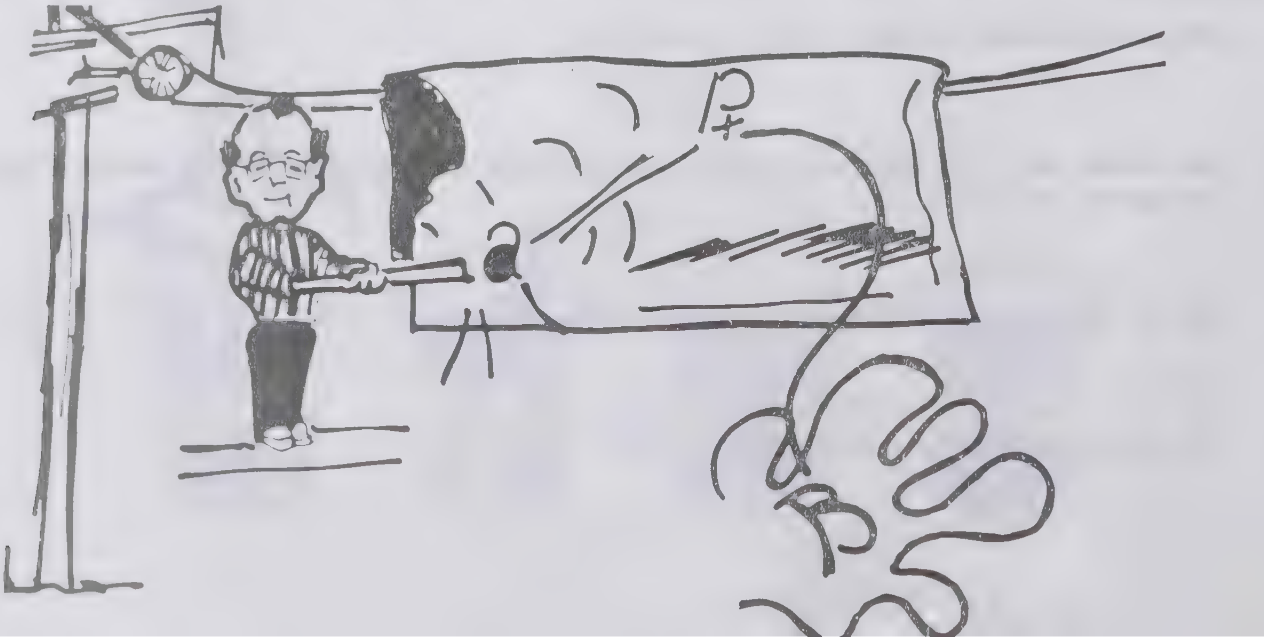
The Bronsted-Lowry theory is often called the *proton transfer* theory because acids and bases are defined according to their effect on hydrogen ions (protons) in solution reactions.

Johannes Nicolaus Bronsted (1879 - 1947)

Bronsted's father (who died when Bronsted was thirteen) was a civil engineer and Bronsted was slated for the same profession. In college, his interest in chemistry made him decide to switch fields. He got his doctorate at the Copenhagen Polytechnic Institute and was selected as a professor of Chemistry at the University of Copenhagen.

Bronsted's work on how acids and bases catalyze reactions forced him to clarify what acids and bases were. The classic definitions were that an acid was a substance that gave up hydrogen ions in solution and bases were substances that gave up hydroxide ions in solution. In 1923 Bronsted proposed that if acids gave up a hydrogen ion, bases took up hydrogen ions in solution and the hydroxide ion was a strong base. (The same theory was proposed at the same time by T.M. Lowry in Britain.) This theory represented a greater flexibility that made it possible to extend acid-base theory to examples that could not be explained before that time.

After World War II, during which Bronsted distinguished himself by his firm anti-Nazi attitude, he was elected to the Danish Parliament but died before he could take his seat.



ACIDS AND BASES

BRONSTED—LOWRY THEORY

Bronsted—Lowry Definitions

An *acid* is a chemical species (anion, cation or molecule) that *loses* a proton in a chemical reaction.

A *base* is a chemical species that *gains* a proton in a chemical reaction.

An *acid-base reaction* is a chemical reaction in which a proton is transferred from an acid to a base to form a new acid and a new base. The new species may transfer the proton back to reform the original species.

General Equation

$$\underset{\text{acid}}{HX} + \underset{\text{base}}{Q} \rightleftharpoons \underset{\text{base}}{X^-} + \underset{\text{acid}}{HQ^+}$$

The above example can be read as: the acid HX loses (donates) a proton (H^+) to the base Q which form the new species X^- and HQ^+. These new species can collide to transfer the proton back again, which means that when the reverse reaction is considered, HQ^+ is an acid and X^- is a base.

Bronsted-Lowry theory requires that the reaction be specified before a substance can be classified as an acid or a base, since many species such as $HCO_3^-{}_{(aq)}$ can either accept or donate protons (depending on the nature of the other reactants present).

Example 1:

$$\underset{\text{base}}{HCO_3^-{}_{(aq)}} + \underset{\text{acid}}{H_3O^+{}_{(aq)}} \rightleftharpoons \underset{\text{acid}}{H_2CO_{3(aq)}} + \underset{\text{base}}{H_2O_{(\ell)}}$$

$HCO_3^-{}_{(aq)}$ is acting as a base.

Example 2:

$$\underset{\text{acid}}{HCO_3^-{}_{(aq)}} + \underset{\text{base}}{S^{2-}{}_{(aq)}} \rightleftharpoons \underset{\text{acid}}{HS^-{}_{(aq)}} + \underset{\text{base}}{CO_3^{2-}{}_{(aq)}}$$

$HCO_3^-{}_{(aq)}$ is acting as an acid.

In the reaction of an acid with water, the H_2O molecule acts as a base since it accepts a proton from the acid.

Example 3:

$$\underset{\text{acid}}{HNO_{2(aq)}} + \underset{\text{base}}{H_2O_{(\ell)}} \rightleftharpoons \underset{\text{acid}}{H_3O^+{}_{(aq)}} + \underset{\text{base}}{NO_2^-{}_{(aq)}}$$

$H_3O^+{}_{(aq)}$ is called the *hydronium ion* and will be discussed at length later in this Unit.

The concept of reaction reversibility is important for the understanding of chemical reactions in general, but particularly important in acid-base reactions. In any chemical reaction, the change of substances can occur *either way*. Equations are often written with a double arrow ($\rightleftharpoons$) to emphasize the reversibility concept.

ACIDS AND BASES

BRONSTED-LOWRY EQUATIONS

Label each species in the following equations as an acid or base.

1. $HSO_3^-{}_{(aq)} + H_2O_{(\ell)} \rightleftharpoons H_3O^+{}_{(aq)} + SO_3^{2-}{}_{(aq)}$

________ ________ ________ ________

2. $NH_{3(aq)} + H_2O_{(\ell)} \rightleftharpoons NH_4^+{}_{(aq)} + OH^-{}_{(aq)}$

________ ________ ________ ________

3. $HF_{(aq)} + HSO_3^-{}_{(aq)} \rightleftharpoons F^-{}_{(aq)} + H_2SO_{3(aq)}$

________ ________ ________ ________

4. $H_2SO_{3(aq)} + HS^-{}_{(aq)} \rightleftharpoons HSO_3^-{}_{(aq)} + H_2S_{(aq)}$

________ ________ ________ ________

Write in the products for each reaction that follows and label each species in each equation as an acid or a base. Assume the transfer of only one proton.

5. $HNO_{3(aq)} + C\ell^-{}_{(aq)} \rightleftharpoons$ ________________ + ________________

________ ________ ________________ ________________

6. $CH_3COOH_{(aq)} + SO_4^{2-}{}_{(aq)} \rightleftharpoons$ ________________ + ________________

________ ________ ________________ ________________

7. $HI_{(aq)} + NO_2^-{}_{(aq)} \rightleftharpoons$ ________________ + ________________

________ ________ ________________ ________________

8. $HCO_3^-{}_{(aq)} + S^{2-}{}_{(aq)} \rightleftharpoons$ ________________ + ________________

________ ________ ________________ ________________

9. $CO_3^{2-}{}_{(aq)} + H_2O_{(\ell)} \rightleftharpoons$ ________________ + ________________

________ ________ ________________ ________________

10. List all species in the nine reactions above that have appeared both as an acid and as a base.

EQUILIBRIUM

Equilibrium

As discussed in *ALCHEM* Unit I an equilibrium is an established state of a closed system Equilibrium is recognized by the lack of visible change, but on a molecular level, two processes are occurring. At equilibrium the rate of the forward process always equals the rate of the reverse process so the system *appears* static.

The *state* of an equilibrium is defined as being the relative proportion of reactants and products that is found at equilibrium. An example of a state of equilibrium would be 75% of the reactants remaining and only 25% of the system being in the form of products. Such a state necessitates the use of equilibrium arrows ($\rightleftharpoons$) shown previously.

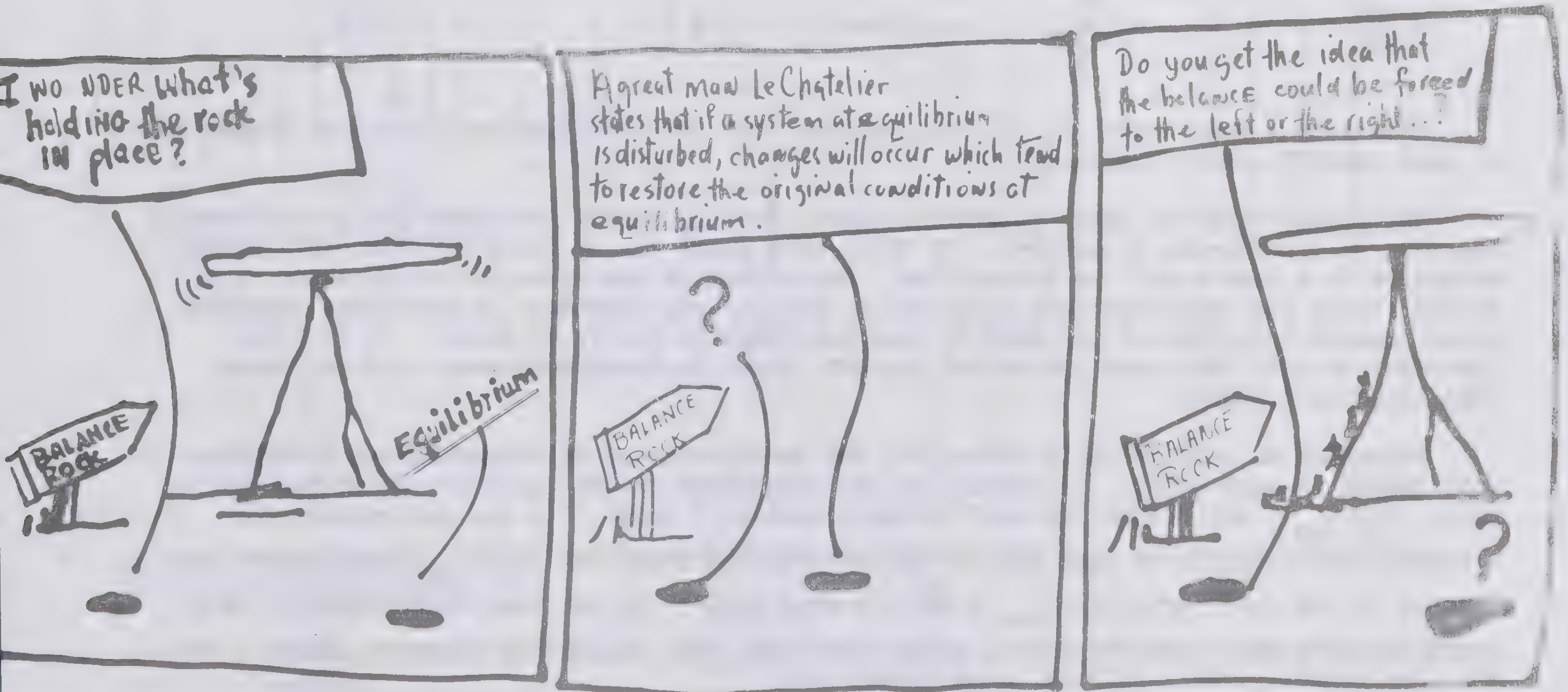
There are some examples of states of equilibrium that *for all practical purposes* are 100% in one direction. Such equilibria are usually indicated by a *one directional arrow*, either $\longrightarrow$ or $\longleftarrow$ depending on whether the forward or reverse reaction predominates. An equilibrium that is essentially 100% in one direction can be thought of as being a stoichiometric reaction in that direction.

The natural state of equilibrium may be altered by imposing a force or change on a system. The effect of an imposed change can be predicted by the *Le Chatelier Principle*. This principle was formulated by Henri Louis Le Chatelier who was a Professor of Chemistry at the University of Paris. The principle states that *if a system at equilibrium is disturbed, changes will occur which tend to restore the original conditions at equilibrium.*

For example, in an equilibrium system

$$\underset{\text{acid}}{HX} + \underset{\text{base}}{Q} \rightleftharpoons \underset{\text{base}}{X^-} + \underset{\text{acid}}{HQ^+}$$

if a large amount of X^- were added to the equilibrium, the system would eliminate *some* of the added X^- by increasing reaction of X^- with HQ^+ (the reverse reaction) to form more reactants. (The system can never *completely* counteract an imposed change.) This reaction is referred to as the equilibrium *shifting to the left*. The result of *any* imposed change on a system at equilibrium can similarly be predicted using Le Chatelier's Principle.


I WONDER what's holding the rock in place?
BALANCE ROCK
Equilibrium
A great man Le Chatelier states that if a system at equilibrium is disturbed, changes will occur which tend to restore the original conditions at equilibrium.
?
BALANCE ROCK
Do you get the idea that the balance could be forced to the left or the right...?
BALANCE ROCK
?

THE SPECIAL NATURE OF H_2O IN BRONSTED-LOWRY THEORY

Since water is almost always the solvent for acid-base reactions, a thorough understanding of the properties of water is essential for understanding acid-base theory. The three areas of most importance are:

1. the nature of the hydronium ion, H_3O^+
2. the dissociation of water
3. the reaction of $H_2O_{(\ell)}$ as a Bronsted-Lowrey acid or base.

The Hydronium Ion

Most chemists believe that since a hydrogen ion (H^+) is the smallest positive ion (due to the absense of *any* electrons) it is strongly attracted to the negative end of polar water molecules. Experimental evidence does not show exactly how many molecules are bonded to a proton in aqueous solution. In fact the number may be variable. Chemists make use of formulas written as $H_3O^+{}_{(aq)}$ and $H^+{}_{(aq)}$. *Both* refer to hydrogen ions in solution, and both formulas are commonly used. The $H_3O^+{}_{(aq)}$ form, called a *hydronium* ion, emphasizes the action of water molecules and the special nature of protons in aqueous solution. *The $H_3O^+{}_{(aq)}$ notation is especially useful in the Bronsted-Lowry theory of acids and bases, and will be used extensively in this unit.*

The Dissociation of Water

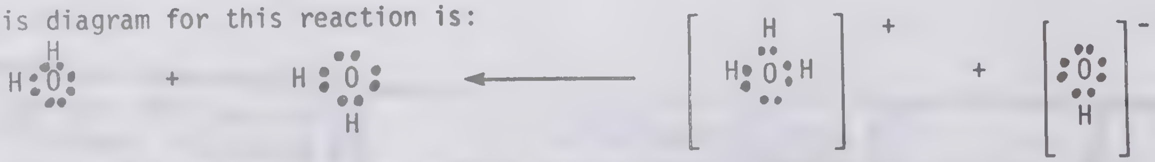
The conductivity of a sample of pure water is negligible in most tests, but can be measured with sensitive equipment. Apparently water molecules dissociate into ions to a very slight extent. The equation for this dissociation can be expressed as:

Equation 1: $H_2O_{(\ell)} \longleftarrow H^+{}_{(aq)} + OH^-{}_{(aq)}$

In Bronsted-Lowry theory this dissociation is thought of as the occasional transfer of a proton from one water molecule to another, expressed as:

Equation 2: $H_2O_{(\ell)} + H_2O_{(\ell)} \longleftarrow H_3O^+{}_{(aq)} + OH^-{}_{(aq)}$

A Lewis diagram for this reaction is:



Equations 1 and 2 refer to the same process but the Bronsted-Lowry equation (Equation 2) is more specific about what occurs.

The dissociation of water as written above, *heavily* favors reactants (is a quantitative reaction in the reverse direction). At 25°C, only about two of every billion (10^9) water molecules in a sample will be dissociated. The extent of the dissociation of water is a trivial thing for many reactions in solution, but is *very important* in acid-base reactions since aqueous solutions of any acid or base are composed mostly of water. It is also important to note that water molecules can act either as Bronsted-Lowrey acids or bases. (See Equation 2 above.)

According to LeChatelier's Principle, the concentrations of hydronium and hydroxide ions depend on each other. An increase in the concentration (at equilibrium) of one species (e.g., $H_3O^+{}_{(aq)}$) will cause the equilibrium (Equation 2 above) to try and restore the original concentration of that species by reacting the added ion ($H_3O^+{}_{(aq)}$) with other ion present in the equilibrium ($OH^-{}_{(aq)}$) to form more water. It has been experimentally determined that the molar concentrations of the two ions, when multiplied together *always* gives the same number, 1.00×10^{-14} at room temperature.

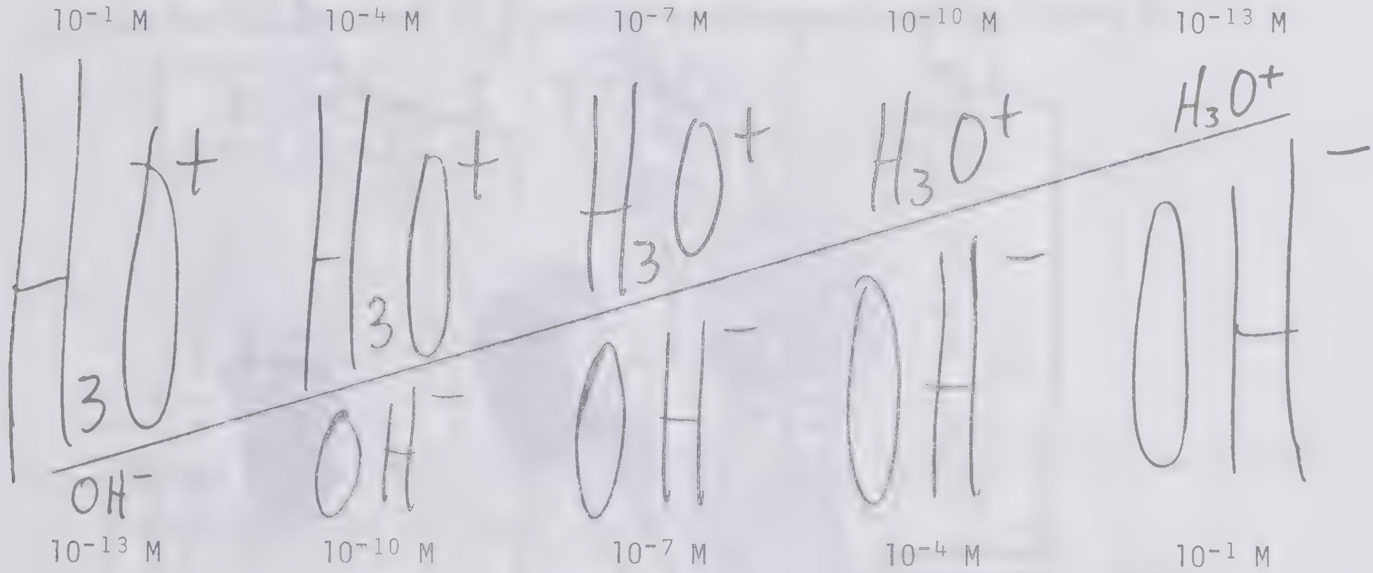
THE SPECIAL NATURE OF H_2O IN BRONSTED-LOWRY THEORY

$[H_3O^+_{(aq)}][OH^-_{(aq)}] = 1.00 \times 10^{-14}$ in acid, base or neutral solutions. In pure water $[H_3O^+_{(aq)}] = [OH^-_{(aq)}]$, and therefore $[H_3O^+_{(aq)}] = 1.00 \times 10^{-7}$ M and $[OH^-_{(aq)}] = 1.00 \times 10^{-7}$ M. (The symbol [ ] is read as (*the*) *molar concentration of* the species shown.)

Acid solutions have a higher concentration of $H_3O^+_{(aq)}$ than pure water. Le Chatelier's Principle states that the increased concentration of hydronium ions (imposed change) will cause some of the excess hydronium ions to react with hydroxide ions, thus decreasing the molar concentration of hydroxide ions. The hydronium ion concentration will remain higher than it was originally (in pure water) since the system cannot counteract all the imposed change. At equilibrium the product $[H_3O^+_{(aq)}]\ [OH^-_{(aq)}]$ is always 1.00×10^{-14} in aqueous solutions. Thus if the hydronium ion concentration is greater than it is in pure water, the hydroxide ion concentration in that sample must be *less* than it is in pure water. Correspondingly, solutions that are more basic than water have an increased concentration of $OH^-_{(aq)}$ resulting in a lower concentration of $H_3O^+_{(aq)}$. The product of the two ion concentrations must still be 1.00×10^{-14}.

Two important considerations result from this characteristic of water. *First*, in aqueous solution an acid can be thought of as any substance that increases $[H_3O^+_{(aq)}]$ and a base as any substance that decreases $[H_3O^+_{(aq)}]$. *Secondly*, in any aqueous solution if the concentration of either $H_3O^+_{(aq)}$ or $OH^-_{(aq)}$ is known, the concentration of the other can be calculated from the expression

$$[H_3O^+]\ [OH^-] = 1.00 \times 10^{-14}$$



1. ________ ________ ________ ________ ________

2. ________ ________ ________ ________ ________

Exercise:

1. Determine and write in the blanks provided the product $[H_3O^+_{(aq)}]\ [OH^-_{(aq)}]$ for each situation in the above figure.
2. Label each situation in the space provided above whether the solution would be acidic, basic or neutral.

THE SPECIAL NATURE OF H_2O IN BRONSTED-LOWRY THEORY

Example 1:

Find $[H_3O^+{}_{(aq)}]$ in a cleaning solution that is 0.00400 M KOH. Since KOH is an ionic compound, it is completely dissociated and

$[OH^-{}_{(aq)}]$ = 0.00400 = 4.00×10^{-3} M.

$$[H_3O^+{}_{(aq)}]\,[OH^-{})_{aq)}] = 1.00 \times 10^{-14}$$

$$[H_3O^+{}_{(aq)} = \frac{1.00 \times 10^{-14}}{[OH^-{}_{(aq)}]}$$

$$= \frac{1.00 \times 10^{-14}}{4.00 \times 10^{-3}\ M}$$

$$= 0.250 \times 10^{-11}\ M$$

$$= \boxed{2.50 \times 10^{-12}\ M}$$

Example 2:

Find $[OH^-{}_{(aq)}]$ of a solution of apple juice in which $[H_3O^+{}_{(aq)}] = 6.00 \times 10^{-3}$ M.

$$[H_3O^+{}_{(aq)}]\,[OH^-{}_{(aq)}] = 1.00 \times 10^{-14}$$

$$[OH^-{}_{(aq)}] = \frac{1.00 \times 10^{-14}}{[H_3O^+{}_{(aq)}]}$$

$$= \frac{1.00 \times 10^{-14}}{6.00 \times 10^{-3}}$$

$$= \boxed{1.67 \times 10^{-12}\ M}$$

CASHIER
What's so special about water?
COMING ATTRACTIONS
blind faith is a political skill

ACIDS AND BASES

THE SPECIAL NATURE OF H_2O IN BRONSTED-LOWRY THEORY

1. Calculate the $[H_3O^+_{(aq)}]$ in milk of magnesia that has $[OH^-] = 1.43 \times 10^{-4}$ M.

[illegible]

2. Calculate the $[OH^-_{(aq)}]$ in lime juice where $[H_3O^+_{(aq)}] = 1.3 \times 10^{-2}$ M.

[illegible]

3. Find the concentration of hydroxide ions in grapefruit juice that has

   $[H_3O^+_{(aq)}] = 2.40 \times 10^{-3}$ M.

[illegible]

4. A recipe for lye soap calls for 3.40 g of NaOH to be dissolved in 15.0 ml of water. Find $[OH^-_{(aq)}]$ and $[H_3O^+_{(aq)}]$ in the solution.

[illegible] mol/L

[illegible] = 5.? mol/L

pH

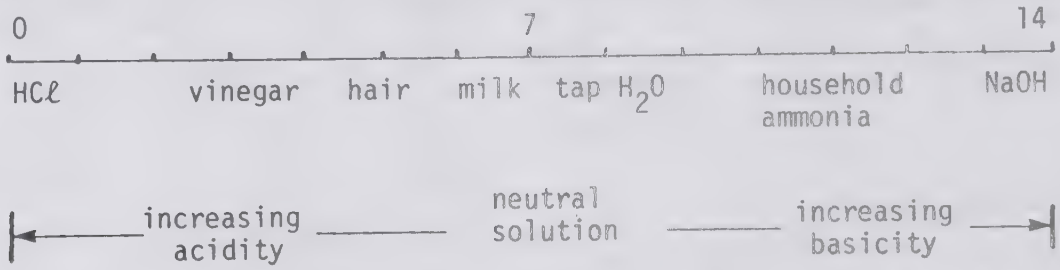
In 1909 a Danish chemist, Soren Sorenson, developed a simplified system for referring to the degree of acidity of a solution. He used the term pH for *potenz* (power) of hydrogen, thus pH refers to the concentration of hydronium (hydrogen) ions in solution. $[H_3O^+_{(aq)}]$ ranges from about 20 M down to 1.0×10^{-15} M in common aqueous solutions. A range on the order of a quadrillion to one can only be expressed by using powers of ten, so pH was based on the following formula:

$$pH = -(\log_{10}[H_3O^+_{(aq)}])$$

For solutions 1.0 M or less in concentration, pH values range from 0.0 to 14.0. Pure water is considered neutral with a pH of 7.0. pH values lower than 7.0 are acidic, while pH values higher than 7.0 are basic.

The pH values of some common solutions are given in the pH scale and Table N1 following

pH Scale



My hair may not be clean but it's Natural!!
Low pH Shampoo

pH

Table N1

Approximate pH of Some Common Substances

Substance	$[H_3O^+_{(aq)}]$ $(\frac{mol}{\ell})$	$[OH^-_{(aq)}]$ $(\frac{mol}{\ell})$	$[H_3O^+_{(aq)}][OH^-_{(aq)}]$	Typical pH	Actual pH Range
1.0 M $HC\ell$	1.0	1.0×10^{-14}	1.0×10^{-14}	0.0	0.0
limes	1.3×10^{-2}	7.7×10^{-13}	1.0×10^{-14}	1.9	1.8 - 2.0
gastric juice	1.0×10^{-2}	1.0×10^{-12}	1.0×10^{-14}	2.0	1.0 - 3.0
lemons	5.0×10^{-3}	2.0×10^{-12}	1.0×10^{-14}	2.3	2.2 - 2.4
vinegar	1.3×10^{-3}	7.7×10^{-12}	1.0×10^{-14}	2.9	2.4 - 3.4
pop	1.0×10^{-3}	1.0×10^{-11}	1.0×10^{-14}	3.0	1.8 - 4.0
rhubarb	8.0×10^{-4}	1.3×10^{-11}	1.0×10^{-14}	3.1	3.1 - 3.2
apples (and cider)	8.0×10^{-4}	1.3×10^{-11}	1.0×10^{-14}	3.1	2.9 - 3.3
wines	4.0×10^{-4}	2.5×10^{-11}	1.0×10^{-14}	3.4	2.8 - 3.8
beers	3.2×10^{-5}	3.2×10^{-10}	1.0×10^{-14}	4.5	4.0 - 5.0
bread (white)	3.1×10^{-6}	3.2×10^{-9}	1.0×10^{-14}	5.5	5.0 - 6.0
asparagus	2.5×10^{-6}	4.0×10^{-9}	1.0×10^{-14}	5.6	5.4 - 5.8
potatoes	1.6×10^{-6}	6.3×10^{-9}	1.0×10^{-14}	5.8	5.6 - 6.0
salmon	6.3×10^{-7}	1.6×10^{-8}	1.0×10^{-14}	6.2	6.1 - 6.3
corn	6.3×10^{-7}	1.6×10^{-8}	1.0×10^{-14}	6.2	6.0 - 6.5
rain water	6.3×10^{-7}	1.6×10^{-8}	1.0×10^{-14}	6.2	4.0 - 6.5
milk (cow)	3.2×10^{-7}	3.2×10^{-8}	1.0×10^{-14}	6.5	6.3 - 6.6
saliva	1.0×10^{-7}	1.0×10^{-7}	1.0×10^{-14}	7.0	6.5 - 7.5
pure water	1.0×10^{-7}	1.0×10^{-7}	1.0×10^{-14}	7.0	7.0
blood	4.0×10^{-8}	2.5×10^{-7}	1.0×10^{-14}	7.4	7.35 - 7.45
swimming pool water	4.0×10^{-8}	2.5×10^{-7}	1.0×10^{-14}	7.4	7.2 - 7.6
eggs (white)	3.0×10^{-8}	3.3×10^{-7}	1.0×10^{-14}	7.8	7.6 - 8.0
sea water	3.2×10^{-9}	3.2×10^{-6}	1.0×10^{-14}	8.5	8.0 - 9.0
household ammonia	1.0×10^{-11}	1.0×10^{-3}	1.0×10^{-14}	11.0	10.6 - 11.6
1.0 M NaOH	1.0×10^{-14}	1.0	1.0×10^{-14}	14.0	14.0

Consult the *CRC Handbook* under *pH* for further pH data.

ACIDS AND BASES

pH OF COMMON SUBSTANCES – DEMO N2

Purpose:

To introduce the pH meter and find the pH of some common substances.

Materials:

- pH meter (or pH paper)
- 50 ml samples of pops, shampoos, fruit juices, water from various sources, etc. that students bring to class.
- several 100 ml beakers

Procedure:

1. Calibrate the pH meter with a solution of known pH.
2. Pour approximately 50 ml of a sample into a clean 100 ml beaker.
3. Immerse the electrodes into the solution and record the pH.
4. Repeat with other samples, rinsing the electrodes with distilled water between pH readings. For shampoos, soaps, etc. it is usually desireable to find the pH of the original substance *and* the pH of a diluted solution of the substance.

Observations:

Questions:

1. What generalizations can be made about the pH of foods?

2. What generalizations can be made about the pH of soaps and other cleaners?

3. What generalizations can be made about the pH of water?

pH CALCULATIONS

Calculations Involving pH

It is apparent from the definition of pH that logarithms are necessary for calculations involving pH. For logarithms, refer to the table of 2 - place logs on the *ALCHEM* data sheet. pH values should go to two decimal places, and concentrations should show two-figure accuracy, when working from this table.

Study the examples and complete the problems that follow.

Example 1:

Calculate the pH of a sample of rain water in which $[H_3O^+_{(aq)}] = 1.00 \times 10^{-4}$ M.

$$\begin{aligned} pH &= -\log [H_3O^+_{(aq)}] \\ &= -\log 1.00 \times 10^{-4} \\ &= -(0.00 + (-4.00)) \\ &= 4.00 \end{aligned}$$

Note:

1. $\log 10^{-x} = -x$
2. $\log xy = \log x + \log y$

Note:

For calculations where the concentration of hydronium ion is 1.0 *times some power of* 10, *the pH is simply the negative of the exponent; e.g., if* $[H_3O^+_{(aq)}] = 1.00 \times 10^{-7}$, *then the pH* = 7.00.

Example 2:

pH from $[H_3O^+_{(aq)}]$.

Find the pH of a wart remover containing nitric acid in which $[H_3O^+_{(aq)}] = 0.25$ M.

Note:

It is convenient to express concentrations in scientific notation in calculations of this type.

$$\begin{aligned} [H_3O^+_{(aq)}] &= 2.5 \times 10^{-1} \text{ M} \\ pH &= -(\log[H_3O^+_{(aq)}]) \\ &= -(\log 2.5 \times 10^{-1}) \\ &= -(0.40 - 1) \\ &= -0.40 + 1 \\ &= 0.60 \end{aligned}$$

Example 3:

pH from $[OH^-_{(aq)}]$

Find the pH of a 0.0060 M KOH solution that is used to remove grease from a garage floor. First, use $[OH^-_{(aq)}]$ to find $[H_3O^+_{(aq)}]$.

$[OH^-_{(aq)}] = 6.0 \times 10^{-3}$ M (assuming complete dissociation)

$$\begin{aligned} [H_3O^+_{(aq)}] &= \frac{1.00 \times 10^{-14}}{[OH^-_{(aq)}]} \\ &= \frac{1.00 \times 10^{-14}}{6.0 \times 10^{-3}} \end{aligned}$$

Example 4:

$[H_3O^+_{(aq)}]$ from pH

Find $(H_3O^+_{(aq)}]$ in a bottle of pop that has a pH of 2.09

First write the pH value as the next largest integer minus a two place decimal; e.g., 2.09 = 3.00 - 0.91 then,

$$\begin{aligned}[H_3O^+_{(aq)}] &= \text{antilog } (-\text{pH}) \\ &= \text{antilog } (-(3.00 - 0.91)) \\ &= \text{antilog } (0.91 - 3.00) \\ &= 8.2 \times 10^{-3}\ \text{M}\end{aligned}$$

Note:
The antilog of a -pH is the number whose log is (-pH). Since the antilog of a negative fraction cannot be obtained directly, it is necessary to rewrite the number as a negative whole number plus a positive fractional number.

Example 5:

$[OH^-_{(aq)}]$ from pH

Find $[OH^-_{(aq)}]$ in a solution with a pH = 2.92.

First, solve for $[H_3O^+_{(aq)}]$.

$$\text{pH} = 2.92 = 3.00 - 0.08$$

$$\begin{aligned}[H_3O^+_{(aq)}] &= \text{antilog } (-\text{pH}) \\ &= \text{antilog } (-(3.00 - 0.08)) \\ &= \text{antilog } (0.08 - 3.00) \\ &= 1.2 \times 10^{-3}\ \text{M}\end{aligned}$$

Then, solve for $[OH^-]$.

$$\begin{aligned}[OH^-_{(aq)}] &= \frac{1.0 \times 10^{-14}}{[H_3O^+_{(aq)}]} \\ &= \frac{1.0 \times 10^{-14}}{1.2 \times 10^{-3}} \\ &= 0.83 \times 10^{-11} \\ &= 8.3 \times 10^{-12}\ \text{M}\end{aligned}$$

Remedial Exercise (Use the 2-place logs on the *ALCHEM* 30 data sheet.)

1. log 100 = ________________
2. log 1.00×10^{-13} = ________________
3. -(log 2.1×10^{-4}) = ________________
4. -(log 0.527) = ________________
5. -(log 4.61×10^{-12}) = ________________
6. antilog 3.00 = ________________
7. antilog 1.85 = ________________
8. antilog (-9.00) = ________________
9. antilog (-2.62) = ________________
10. antilog (-7.43) = ________________

ACIDS AND BASES

pH CALCULATIONS

Questions:

1. Find the pH of a lime that has $[H_3O^+_{(aq)}]$ = 0.0120 M.

2. Find $[H_3O^+_{(aq)}]$ and pH of a solution made by dissolving 10.0 g of KOH in water to make 4.00 ℓ of solution.

3. A sample of sea water was tested and found to have a pH of 8.65.
Find $[H_3O^+_{(aq)}]$ and $[OH^-_{(aq)}]$.

4. Consider a sample of stomach contents which has a pH of 1.40.
Find $[H_3O^+_{(aq)}]$ and $[OH^-_{(aq)}]$.

pH CALCULATIONS

	Solution (acidic or basic)	$[H_3O^+_{(aq)}]$	$[OH^-_{(aq)}]$	pH
5.		4.0×10^{-6} M		
6.				9.50
7.			2.0×10^{-11} M	
8.		10 M		
9.				15.00
10.			4.4×10^{-2} M	

Work Space

pH OF ACIDS AND BASE - DEMO N3

Purpose:

To determine that the pH of 0.10 M solutions of some common acids and bases varies.

Predemo Information:

A *strong acid* is defined as one that completely reacts to form hydronium ions in aqueous solution. In the following demo, the same concentration of all acids and bases is used. If an acid is a strong acid, there should be complete reaction to $H_3O^+_{(aq)}$ and thus a low pH.

If an acid is a weak acid, few molecules should react with water and therefore a pH closer to 7 would be expected. The strongest base should have the highest pH since a strong base is defined as one that readily reacts with hydronium ions.

Materials:

- pH meter
1 - 250 ml beaker
- wash bottle containing distilled water
9 - 150 ml beakers containing about 100 ml of each of the following 0.10 M aqueous solution
  - hydrochloric acid, $HC\ell_{(aq)}$
  - nitric acid, $HNO_{3(aq)}$
  - oxalic acid, $HOOCCOOH_{(aq)}$
  - sodium hydrogen sulfate, $NaHSO_{4(aq)}$
  - acetic acid, $CH_3COOH_{(aq)}$
  - ammonium chloride, $NH_4C\ell_{(aq)}$
  - potassium hydroxide, $KOH_{(aq)}$
  - sodium hydroxide, $NaOH_{(aq)}$
  - ammonia, $NH_{3(aq)}$

Procedure:

1. Calibrate the pH meter with a solution of known pH.
2. Test each solution by dipping the electrodes of the pH meter into the solution and record the pH. Rinse the electrodes with distilled water before dipping them into the next solution.

Observations:

Formula of Substance	pH	Order 1,2,3	Order From ALCHEM Data Sheet
$HC\ell_{(aq)}$			
$HNO_{3(aq)}$			
$HOOCCOOH_{(aq)}$			
$NaHSO_{4(aq)}$			
$CH_3COOH_{(aq)}$			
$NH_4C\ell_{(aq)}$			
$KOH_{(aq)}$			
$NaOH_{(aq)}$			
$NH_{3(aq)}$			

Questions:

1. Number the acid solutions 1 to 6 in order of increasing pH.
2. Number the base solutions 1 to 3 in order of decreasing pH.
3. Compare the order of pH obtained for the acids with the order of strength from the *ALCHEM* data sheet.
4. Explain why the pH of the acid solutions varied although the concentrations of the solutions were identical.

STRENGTH VERSUS CONCENTRATION FOR ACIDS AND BASES

Definitions

1. *Solubility* - the quantity of a substance that will dissolve in a solvent to form a saturated solution at a given temperature (maximum concentration)

2. *% Reaction* - the percentage of a dissolved acid that reacts to form hydronium ions in water

The demonstration of pH, Demo N3, illustrates the differences in the two concepts of strength and concentration. Chemists routinely refer to *strong* and *weak* acids and bases. The terms strong and weak refer to the degree of reaction of the substances, not their concentration in solution.

Strength of Acids

A *strong acid* is an acid that *reacts* completely when forming an aqueous solution. $HC\ell_{(aq)}$ is the most common example of a strong acid. When $HC\ell$ molecules dissolve in water they react completely, forming as many hydronium ions as possible. (Note single arrow in the equation below.)

$$HC\ell_{(g)} + H_2O_{(\ell)} \longrightarrow H_3O^+_{(aq)} + C\ell^-_{(aq)}$$

In a solution of a strong acid, $[H_3O^+]$ *equals* the concentration of the solution; i.e., if a 0.10 M solution of a strong acid is prepared, the $[H_3O^+_{(aq)}]$ of that solution will be 0.10 M.

In *ALCHEM* Unit M, reduction potentials were obtained by comparing all oxidizing agents to the hydrogen half cell. Similarly in the *ALCHEM* strength of acid table, it is necessary to compare the ability of the *same* base to remove protons from acids. For Bronsted-Lowry proton transfer reactions, water was chosen as the base used for comparison.

A *weak acid* is an acid that reacts only slightly in aqueous solution. Acetic acid is an example of a common weak acid. When CH_3COOH molecules dissolve in water, only a small percentage react (only 0.4% in 1.0 M solution or 4 molecules of every 1000 that dissolve). Thus for weak acids $[H_3O^+_{(aq)}]$ is much less than the concentration of the solution.

$$CH_3COOH_{(aq)} + H_2O_{(\ell)} \overset{\longleftarrow}{\rightleftharpoons} H_3O^+_{(aq)} + CH_3COO^-_{(aq)}$$

Notice that in the weak acid reaction shown above, equilibrium arrows are used to indicate that the reaction is not complete in either direction. In *ALCHEM*, wherever it is necessary to indicate the state of equilibrium, a single arrow will be placed above the equilibrium arrows to show whether the forward or reverse reaction predominates. The term *predominates* refers to a *greater tendency to occur*. The greater tendency to occur results *from* the different proton attracting ability of the two bases involved in the reaction. (This idea is expanded on later in this Unit; Page N25.) The greater tendency of either the forward or reverse reaction to occur results *in* either a greater concentration of products or reactants (respectively) at equilibrium. The percent reaction of acids with water is used on the following two pages to illustrate the idea of either the forward or reverse reaction with water predominating.

Remember, the strength of an acid refers to the ease with which it will *lose* protons. It is quite likely that a dilute solution of a strong acid will provide a higher hydronium ion concentration that a concentrated solution of a weak acid since so few of the weak acid molecules react with water.

STRENGTH VERSUS CONCENTRATION FOR ACIDS AND BASES

Weak Acids and pH

When calculating $[H_3O^+]$ or pH of a 1.0 M solution of a weak acid it is necessary to consider the degree to which the acid reacts with water.

Example:

Oxalic acid is used to remove rust from stained bathroom fixtures in areas where the water supply contains large amounts of iron.

Calculate the pH of a 1.00 M solution of oxalic acid.

$$H_2O_{(\ell)} + HOOCCOOH_{(aq)} \rightleftharpoons H_3O^+_{(aq)} + HOOCCOO^-_{(aq)}$$

From the *ALCHEM* data sheet, the reaction occurs to the extent of 21%.

$$\begin{aligned}[H_3O^+_{(aq)}] &= 21\% \text{ of the concentration of the acid solution}\\ &= \frac{21}{100} \times 1.00\text{ M} = 0.21\text{ M}\\ \text{pH} &= -\log\ 2.1 \times 10^{-1}\text{ M}\\ &= -(0.32 - (-1))\\ &= -(-0.68)\\ &= 0.68\end{aligned}$$

Notes:

1. Calculating $[H_3O^+_{(aq)}]$ from % dissociation on the *ALCHEM* data sheet is only valid for a 1.0 M solution since % dissociation varies with concentration.

2. Calculating $[H_3O^+]$ by considering % dissociation of an ion containing hydrogen (e.g., HCO_3^-, $HOOCCOO^-$) is not valid since these ions can act as either acids or bases and with water often react as bases.

Questions:

1. Hydrofluoric acid is often used to etch glass since it is one of the few substances that reacts with glass. Calculate the $[H_3O^+_{(aq)}]$, $[OH^-_{(aq)}]$ and pH of a 1.0 M solution of $HF_{(aq)}$.

2. Calculate the $[H_3O^+_{(aq)}]$ and pH of a vinegar sample (1.00 M solution of acetic acid).

STRENGTH VERSUS CONCENTRATION FOR ACIDS AND BASES

3. Calculate the $[H_3O^+_{(aq)}]$ and pH of a 1.00×10^{-2} M solution of hydrobromic acid.

4. Compare the $[H_3O^+_{(aq)}]$ in the dilute solution of the strong acid (Question 3) with the $[H_3O^+_{(aq)}]$ in the more concentrated solution of the weak acid (Question 2).

5. Boric acid is a weak acid that is often used as an eyewash. A 0.100 M solution of $H_3BO_{3(aq)}$ has a pH of 5.2?

$$H_2O_{(\ell)} + H_3BO_{3(aq)} \overset{?}{\rightleftharpoons} H_3O^+_{(aq)} + H_2BO_3^-{}_{(aq)}$$

Calculate the % reaction of $H_3BO_{3(aq)}$ with water.

ACIDS AND BASES

PREDICTING REACTIONS FROM BRONSTED - LOWRY THEORY

Bronsted-Lowry acid-base reactions involve proton transfer *to the strongest base from the strongest acid*. (Note the similarity to the process of an electrochemical reaction transferring an electron from the strongest oxidizing agent to the strongest reducing agent.)

The *ALCHEM* acid-base table can be used to predict the products for any Bronsted-Lowry acid - base reaction. In addition, it enables prediction as to whether equilibrium favors reactant or product species. The procedure for predicting products and equilibrium for reactions of this type is as follows.

Step 1: Identify all reactants and spectator ions. Write down the formulas for all possible Bronsted-Lowry acids and bases.

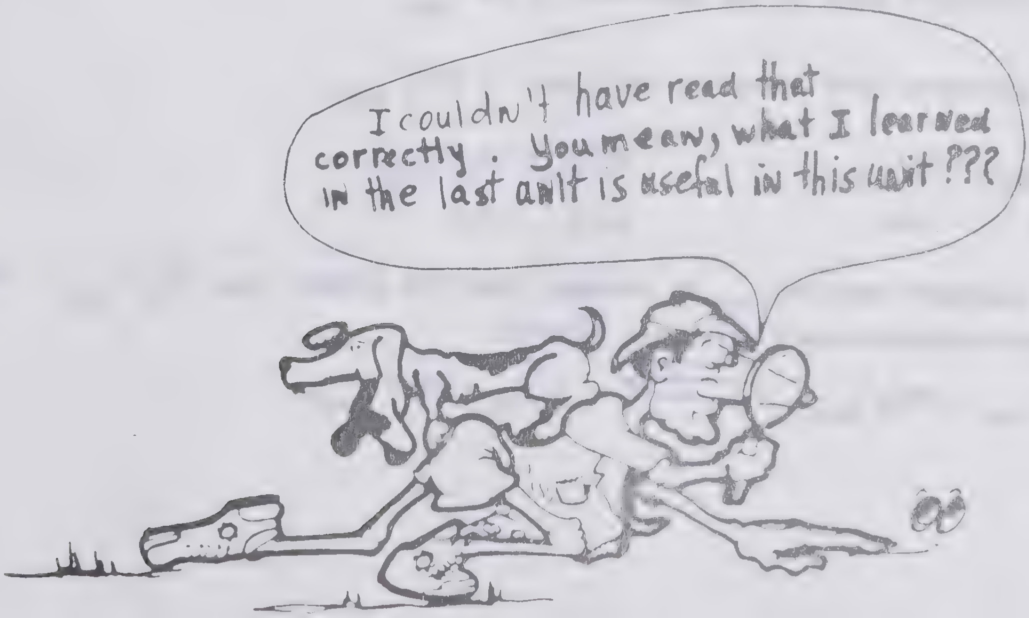
Step 2: Identify the strongest base that exists among the possible reactants. This species will be the proton acceptor.

Step 3: Identify the strongest acid that exists among the *remaining* reactants. This species will be the proton donor.

Step 4: Complete the equation by writing the formulas for the products assuming *one* proton transfers from the acid to the base. Identify each product as an acid or a base.

Step 5: Decide whether equilibrium favors reactants or products. Compare the strengths of the reactant base and the product base. The stronger base attracts protons better than the weaker base and therefore proton transfer to the stronger base predominates. Thus if the product base is the stronger, the reaction of the species on the product side of the equation to form reactants will occur to the greatest extent. This is usually stated as *reverse reaction favored* or *equilibrium favors reactants*. (Note that comparing base strengths to determine predominance in an acid-base reaction is an identical procedure to comparing oxidizing agent strengths to determine spontaneity in an electrochemical reaction.)

In the examples that follow, the state of the equilibrium is indicated by the method stated earlier in the unit. An arrow *above* the equilibrium arrows indicates that reaction in the direction of the arrow predominated (had a greater tendency to occur). If the extent of reaction is essentially 100%, a single arrow in the direction of reaction *replaces* the equilibrium arrows. *It can be assumed that 100% reaction occurs when* $H_3O^+_{(aq)}$ *reacts with any base below* $NO_2^-{}_{(aq)}$ *on the ALCHEM data table or when* $OH^-_{(aq)}$ *reacts with any acid above* HSO_3^- *on the table.*


I couldn't have read that correctly. You mean, what I learned in the last unit is useful in this unit???

ACIDS AND BASES

PREDICTING REACTIONS FROM BRONSTED LOWRY THEORY

Example 1:

Hydrochloric acid is added to a solution of sodium benzoate, NaC_6H_5COO.

Step 1: Hydrochloric acid reacts completely with water and is written as $H_3O^+_{(aq)}$ and $C\ell^-_{(aq)}$. Sodium benzoate is an ionic compound and will completely dissociate into $Na^+_{(aq)}$ and $C_6H_5COO^-_{(aq)}$. Neither $Na^+_{(aq)}$ nor $C\ell^-_{(aq)}$ can act as acids or bases so they must be spectator ions. Water can act as either an acid or a base in every reaction.

possible bases – $C_6H_5COO^-_{(aq)}$ or $H_2O_{(\ell)}$

possible acids – $H_3O^+_{(aq)}$ or $H_2O_{(\ell)}$

Step 2: The strongest base present is $C_6H_5COO^-_{(aq)}$, so it will be the proton acceptor.

Step 3: $H_3O^+_{(aq)}$ is the strongest acid present, so it will be the proton donor.

Step 4:

$$\underset{\text{base}}{C_6H_5COO^-_{(aq)}} + \underset{\text{acid}}{H_3O^+_{(aq)}} \qquad \underset{\text{acid}}{C_6H_5COOH_{(aq)}} + \underset{\text{base}}{H_2O_{(\ell)}}$$

Step 5: The bases competing for protons are $C_6H_5COO^-_{(aq)}$ (forward reaction) and $H_2O_{(\ell)}$ (reverse reaction). According to the *ALCHEM* data sheet, $C_6H_5COO^-_{(aq)}$ is a stronger base and therefore the forward reaction was favored before equilibrium was established. Since this example is a reaction of $H_3O^+_{(aq)}$ with a base below $NO_2^-_{(aq)}$, it can be considered essentially quantitative and a single arrow can be used.

$$C_6H_5COO^-_{(aq)} + H_3O^+_{(aq)} \longrightarrow C_6H_5COOH_{(aq)} + H_2O_{(\ell)}$$

Example 2:

A lye solution is reacted with sodium bicarbonate.

Step 1: The possible reactant acids are $H_2O_{(\ell)}$ and $HCO_3^-_{(aq)}$. The possible reactant bases are $OH^-_{(aq)}$, $H_2O_{(\ell)}$ and $HCO_3^-_{(aq)}$. $Na^+_{(aq)}$ is a spectator ion.

Step 2: The strongest base of the reactants is $OH^-_{(aq)}$.

Step 3: The strongest acid among the reactants is $HCO_3^-_{(aq)}$.

Step 4:

$$\underset{\text{base}}{OH^-_{(aq)}} + \underset{\text{acid}}{HCO_3^-_{(aq)}} \qquad \underset{\text{acid}}{H_2O_{(\ell)}} + \underset{\text{base}}{CO_3^{2-}_{(aq)}}$$

Step 5: The reactant base $OH^-_{(aq)}$ is stronger than the product base $CO_3^{2-}_{(aq)}$. Therefore the forward reaction would have been favored.

$$OH^-_{(aq)} + HCO_3^-_{(aq)} \rightleftharpoons H_2O_{(\ell)} + CO_3^{2-}_{(aq)}$$

ACIDS AND BASES

PREDICTING REACTIONS FROM BRONSTED – LOWRY THEORY

Bronsted-Lowry acid-base reactions involve proton transfer *to the strongest base from the strongest acid.* (Note the similarity to the process of an electrochemical reaction transferring an electron from the strongest oxidizing agent to the strongest reducing agent.)

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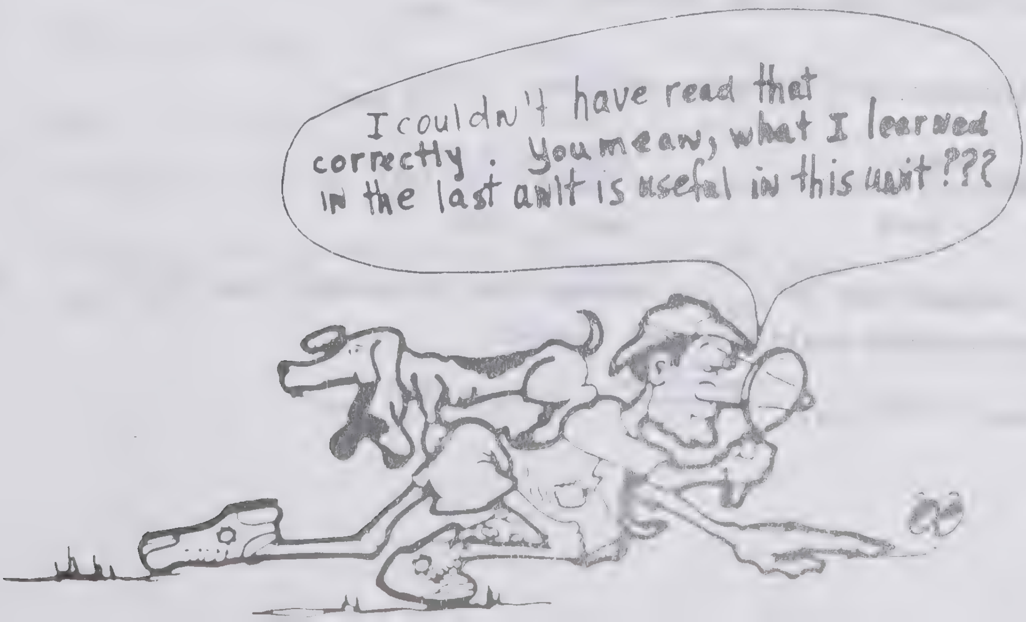
Step 3: Identify the strongest acid that exists among the *remaining* reactants. This species will be the proton donor.

Step 4: Complete the equation by writing the formulas for the products assuming *one* proton transfers from the acid to the base. Identify each product as an acid or a base.

Step 5: Decide whether equilibrium favors reactants or products. Compare the strengths of the reactant base and the product base. The stronger base attracts protons better than the weaker base and therefore proton transfer to the stronger base predominates. Thus if the product base is the stronger, the reaction of the species on the product side of the equation to form reactants will occur to the greatest extent. This is usually stated as *reverse reaction favored* or *equilibrium favors reactants.* (Note that comparing base strengths to determine predominance in an acid-base reaction is an identical procedure to comparing oxidizing agent strengths to determine spontaneity in an electrochemical reaction.)

In the examples that follow, the state of the equilibrium is indicated by the method stated earlier in the unit. An arrow *above* the equilibrium arrows indicates that reaction in the direction of the arrow predominated (had a greater tendency to occur). If the extent of reaction is essentially 100%, a single arrow in the direction of reaction *replaces* the equilibrium arrows. *It can be assumed that 100% reaction occurs when* $H_3O^+{}_{(aq)}$ *reacts with any base below* $NO_2^-{}_{(aq)}$ *on the ALCHEM data table or when* $OH^-{}_{(aq)}$ *reacts with any acid above* HSO_3^- *on the table.*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ACIDS AND BASES

PREDICTING REACTIONS FROM BRONSTED LOWRY THEORY

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possible bases – $C_6H_5COO^-{}_{(aq)}$ or $H_2O_{(\ell)}$

possible acids – $H_3O^+{}_{(aq)}$ or $H_2O_{(\ell)}$

Step 2: The strongest base present is $C_6H_5COO^-{}_{(aq)}$, so it will be the proton acceptor.

Step 3: $H_3O^+{}_{(aq)}$ is the strongest acid present, so it will be the proton donor.

Step 4:

$$\underset{\text{base}}{C_6H_5COO^-{}_{(aq)}} + \underset{\text{acid}}{H_3O^+{}_{(aq)}} \qquad \underset{\text{acid}}{C_6H_5COOH_{(aq)}} + \underset{\text{base}}{H_2O_{(\ell)}}$$

Step 5: The bases competing for protons are $C_6H_5COO^-{}_{(aq)}$ (forward reaction) and $H_2O_{(\ell)}$ (reverse reaction). According to the *ALCHEM* data sheet, $C_6H_5COO^-{}_{(aq)}$ is a stronger base and therefore the forward reaction was favored before equilibrium was established. Since this example is a reaction of $H_3O^+{}_{(aq)}$ with a base below $NO_2^-{}_{(aq)}$, it can be considered essentially quantitative and a single arrow can be used.

$$C_6H_5COO^-{}_{(aq)} + H_3O^+{}_{(aq)} \longrightarrow C_6H_5COOH_{(aq)} + H_2O_{(\ell)}$$

Example 2:

A lye solution is reacted with sodium bicarbonate.

Step 1: The possible reactant acids are $H_2O_{(\ell)}$ and $HCO_3^-{}_{(aq)}$. The possible reactant bases are $OH^-{}_{(aq)}$, $H_2O_{(\ell)}$ and $HCO_3^-{}_{(aq)}$. $Na^+{}_{(aq)}$ is a spectator ion.

Step 2: The strongest base of the reactants is $OH^-{}_{(aq)}$.

Step 3: The strongest acid among the reactants is $HCO_3^-{}_{(aq)}$.

Step 4:

$$\underset{\text{base}}{OH^-{}_{(aq)}} + \underset{\text{acid}}{HCO_3^-{}_{(aq)}} \qquad \underset{\text{acid}}{H_2O_{(\ell)}} + \underset{\text{base}}{CO_3^{2-}{}_{(aq)}}$$

Step 5: The reactant base $OH^-{}_{(aq)}$ is stronger than the product base $CO_3^{2-}{}_{(aq)}$. Therefore the forward reaction would have been favored.

$$OH^-{}_{(aq)} + HCO_3^-{}_{(aq)} \rightleftharpoons H_2O_{(\ell)} + CO_3^{2-}{}_{(aq)}$$

ACIDS AND BASES

PREDICTING REACTIONS FROM BRONSTED-LOWRY THEORY

Example 3:

Ammonia is dissolved in water to make a household cleaner.

Step 1:

The only acid is $H_2O_{(\ell)}$

The bases are $H_2O_{(\ell)}$ and $NH_{3(aq)}$.

Step 2:

$NH_{3(aq)}$ is the stronger base and therefore the proton acceptor.

Step 3:

$H_2O_{(\ell)}$ is the proton donor

Step 4:

$$\underset{\text{acid}}{H_2O_{(\ell)}} + \underset{\text{base}}{NH_{3(aq)}} \qquad\qquad \underset{\text{acid}}{NH_4{}^+{}_{(aq)}} + \underset{\text{base}}{OH^-{}_{(aq)}}$$

Step 5:

The product base $OH^-{}_{(aq)}$ is stronger than the reactant base $NH_{3(aq)}$ and therefore the reverse reaction is favored.

$$H_2O_{(\ell)} + NH_{3(aq)} \overset{\longleftarrow}{\rightleftharpoons} NH_4{}^+{}_{(aq)} + OH^-{}_{(aq)}$$

Note that ammonia is the only example thus far of a molecular compound that dissolves in water to form a basic solution.

Example 4:

Sodium bicarbonate dissolves in water as part of a baking process.

Step 1:

$HCO_3{}^-{}_{(aq)}$ or $H_2O_{(\ell)}$ could be acids.

$HCO_3{}^-{}_{(aq)}$ or $H_2O_{(\ell)}$ could be bases.

Na^+ is a spectator ion.

Step 2:

The strongest base is the reactant $HCO_3{}^-{}_{(aq)}$.

Step 3:

Since $HCO_3{}^-$ will be acting as a base, only $H_2O_{(\ell)}$ can be considered when choosing the strongest acid.

Step 4:

$$\underset{\text{base}}{HCO_3{}^-{}_{(aq)}} + \underset{\text{acid}}{H_2O_{(\ell)}} \qquad\qquad \underset{\text{base}}{OH^-{}_{(aq)}} + \underset{\text{acid}}{H_2CO_{3(aq)}}$$

Step 5:

The product base $OH^-{}_{(aq)}$ is stronger than the reactant base $HCO_3{}^-{}_{(aq)}$.

$$HCO_3{}^-{}_{(aq)} + H_2O_{(\ell)} \longleftarrow OH^-{}_{(aq)} + H_2CO_{3(aq)}$$

ACIDS AND BASES

WRITING ACID—BASE REACTIONS

For each of the following problems, write the reaction equation, label each species as a Bronsted—Lowry acid or base, and indicate whether the forward or reverse reaction was favored before equilibrium (by writing a single arrow over the equilibrium arrows, as in the previous examples.) Assume only one proton is transferred.

1. Solution of $Na_2SO_{3(aq)}$ and $HF_{(aq)}$ are mixed in a beaker. Write the equation for the reaction that occurs, preventing the HF from etching the beaker.

2. A solution of the fertilizer NH_4NO_3 and a solution of $NaCH_3COO$ are mixed.

3. Sodium benzoate is often used as a preservative. Write the equation for solid NaC_6H_5COO dissolving in a solution of $NaHSO_4$.

4. A household ammonia solution is mixed with a solution of nitrous acid.

5. Nitric acid and potassium hydroxide solutions are mixed.

6. Sodium oxide reacts violently with water.

ACIDS AND BASES

WRITING ACID-BASES REACTIONS

7. Sodium sulfite is dissolved into a solution of sulfurous acid.

8. Ammonium fluoride is dissolved in water.

9. Benzoic acid can be used as a starting material for the synthesis of aspirin. Write an equation for the dissolving of benzoic acid in water.

10. In solution, the poisonous gas H_2S (that is sometimes found in crude oil) reacts with carbonate ions.

11. Rain water in an area near a smelter could be a solution of H_2SO_3. Write the equation that represents the reaction of this acid solution with sodium sulfate.

12. Sulfur dioxide from burning coal is reacted with an aqueous solution of ammonia.

13. Sodium bicarbonate may be used directly or in *gripe water* to neutralize excess stomach acid ($HC\ell_{(aq)}$).

BRONSTED–LOWRY REACTIONS - LAB N1

Purpose:

To illustrate that the results of an acid–base reaction can be predicted by Bronsted–Lowry theory.

Prelab Demo Exercise:

Predict the products of the reaction between solutions of sodium hydrogen sulfide and hydrochloric acid.

Prelab Demo Materials:

- stock solution of 1.0 M sodium hydrogen sulfide, NaHS
- stock solution of 1.0 M hydrochloric acid, $HC\ell_{(aq)}$
- 50 ml graduated cylinder
- 250 ml beaker

Prelab Demo Procedure:

1. Using a graduated cylinder, pour 50 ml of 1.0 M sodium hydrogen sulfide solution into a 250 ml beaker.
2. Using a graduated cylinder, add 50 ml of 1.0 M hydrochloric and solution to the solution in the beaker.
3. Identify the product by the odor formed and verify the prediction made in the Prelab Demo Exercise

Prelab Exercises:

Predict the products of the following reactions before doing Part A.

1. An ammonium chloride solution is mixed with a sodium hydroxide solution.

2. Aqueous solutions of sodium acetate and hydrochloric acid are mixed.

3. A sodium hydrogen carbonate solution reacts with a sodium hydrogen sulfate solution.

BRONSTED—LOWRY REACTIONS – LAB N1

Before doing Part B, predict the products of the following reactions *and* whether the forward or reverse reaction is favored as equilibrium is being approached.

4. A methyl orange solution is put into a hydrochloric acid solution.

5: A bromthymol blue solution is mixed with a sodium hydrogen carbonate solution.

6. A phenolphthalein solution is added to a NaOH solution.

Materials:

Part A:

10 ml of 0.10 M solutions of each of the following:

- ammonium chloride, $NH_4C\ell$
- sodium hydroxide, NaOH
- sodium acetate, $NaCH_3COO$
- hydrochloric acid, $HC\ell_{(aq)}$
- sodium hydrogen carbonate, $NaHCO_3$
- sodium hydrogen sulfate, $NaHSO_4$

1 - 10 ml graduated cylinder or pipet

3 - 18 x 100 mm test tubes

Part B:

dropper bottles of the following weak acid solutions:

- methyl orange
- bromthyml blue
- phenolphthalein

1 - 10 ml graduated cylinder or pipet

10 ml of 0.10 M solutions of each of the following;

- hydrochloric acid, $HC\ell_{(aq)}$
- sodium hydrogen carbonate, $NaHCO_3$
- sodium hydroxide, NaOH

3 - 18 x 100 mm test tubes

Procedure:

Part A:

1. Using a graduated cylinder or pipet, place 10 ml of 0.10 M ammonium chloride solution into an 18 x 100 mm test tube.

2. Using a graduated cylinder or pipet, add 10 ml of 0.10 M sodium hydroxide solution to the test tube. Note any observations.

3. Repeat Steps 1 and 2 using a sodium acetate solution and hydrochloric acid.

4. Repeat Steps 1 and 2 using a sodium hydrogen carbonate solution and a sodium hydrogen sulfate solution.

ACIDS AND BASES

BRONSTED—LOWRY REACTIONS - LAB N1

Question:

Were the predictions of products and state of equilibrium made in the Prelab Exercises verified? Support the answers.

Procedure:

Part B:

1. Using a graduated cylinder or pipet, place 10 ml of 0.10 M hydrochloric acid into a 18 x 100 mm test tube.
2. Add a few drops of methyl orange solution to the test tube. Note any observations.
3. Repeat Steps 1 and 2 using a sodium hydrogen carbonate solution and bromthymol blue.
4. Repeat Steps 1 and 2 using a sodium hydroxide solution and phenolphthalein.

Question:

Were the redictions of products and state of equilibrium made in the Prelab Exercises verified? Support the answers.

POLYPROTIC ACIDS AND POLYBASIC SPECIES

Reactions of Polyprotic Acids With Water

Many molecules have more than one proton which can be released to a base. Such acids are called *polyprotic* and include species such as H_2SO_4, H_2CO_3, HOOCCOOH, and H_3PO_4. As a general rule, the first hydrogen (proton) can be removed much more easily than the second, which in turn can be removed more easily than the third (if any). For example $H_2SO_{4(aq)}$ reacts completely with H_2O to form $HSO_4{}^-{}_{(aq)}$ and $H_3O^+{}_{(aq)}$. Then the $HSO_4{}^-{}_{(aq)}$ formed in this step reacts further, $HSO_4{}^-{}_{(aq)} + H_2O_{(\ell)} \rightleftharpoons H_3O^+{}_{(aq)} + SO_4{}^{2}{}_{(aq)}$, but only *partially* (about 10% complete).

Reactions of Polyprotic Acids With Bases Stronger Than Water

In reactions involving polyprotic acids, it is sometimes important to realize that two or three steps may occur.

Example:

If a solution containing $OH^-{}_{(aq)}$ is added to the oxalic acid in rhubarb leaves, the following reactions occur.

Reaction 1:

Step 1: The reactant acid could be $HOOCCOOH_{(aq)}$ or $H_2O_{(\ell)}$.
The reactant base could be $OH^-{}_{(aq)}$ or $H_2O_{(\ell)}$.

Step 2: The strongest base is $OH^-{}_{(aq)}$.

Step 3: The strongest acid is $HOOCCOOH_{(aq)}$.

Step 4: $\underset{\text{base}}{OH^-{}_{(aq)}} + \underset{\text{acid}}{HOOCCOOH_{(aq)}} \qquad \underset{\text{acid}}{H_2O_{(\ell)}} + \underset{\text{base}}{HOOCCOO^-{}_{(aq)}}$

Step 5: $OH^-{}_{(aq)}$ is a much stronger base than $HOOCCOO^-{}_{(aq)}$. The previously mentioned rule applies (reactions between $OH^-{}_{(aq)}$ and any acid above $HSO_3{}^-{}_{(aq)}$ is considered quantitative).

$$OH^-{}_{(aq)} + HOOCCOOH_{(aq)} \longrightarrow H_2O_{(\ell)} + HOOCCOO^-{}_{(aq)}$$

Reaction 2:

$HOOCCOO^-{}_{(aq)}$, the product of the first reaction is still an acid and if more $OH^-{}_{(aq)}$ is added, the $HOOCCOO^-{}_{(aq)}$ can react further.

Step 1: The reactant base could be $OH^-{}_{(aq)}$, $H_2O_{(\ell)}$ or $HOOCCOO^-{}_{(aq)}$.

The reactant acid could be $HOOCCOO^-{}_{(aq)}$ or $H_2O_{(\ell)}$.

Step 2: The strongest reactant base is $OH^-{}_{(aq)}$.

Step 3: The strongest reactant acid is $HOOCCOO^-{}_{(aq)}$.

Step 4: $\underset{\text{base}}{OH^-{}_{(aq)}} + \underset{\text{acid}}{HOOCCOO^-{}_{(aq)}} \qquad \underset{\text{acid}}{H_2O_{(\ell)}} + \underset{\text{base}}{OOCCOO^{2-}{}_{(aq)}}$

Step 5: $OH^-{}_{(aq)} + HOOCCOO^-{}_{(aq)} \longrightarrow H_2O_{(\ell)} + OOCCOO^{2-}{}_{(aq)}$
(according to the previously discussed rule)

The two reactions can be added to give the overall reaction.

$$HOOCCOOH_{(aq)} + 2\ OH^-{}_{(aq)} \longrightarrow 2\ H_2O_{(\ell)} + OOCCOO^{2-}{}_{(aq)}$$

POLYPROTIC ACIDS AND POLYBASIC SPECIES

Reactions of Polybasic Species

Polybasic species are bases that may gain more than one proton.

Example:

If a solution of sodium carbonate is reacted with a strong acid such as $HNO_{3(aq)}$ the reactions are:

Reaction 1:

Step 1: acids are $H_2O_{(\ell)}$ and $H_3O^+{}_{(aq)}$.
bases are $CO_3{}^{2-}{}_{(aq)}$ and $H_2O_{(\ell)}$.
spectator ions are $NO_3{}^-{}_{(aq)}$ and $Na^+{}_{(aq)}$.

Step 2: strongest base is $CO_3{}^{2-}{}_{(aq)}$

Step 3: strongest acid is $H_3O^+{}_{(aq)}$

Step 4: $H_3O^+{}_{(aq)} + CO_3{}^{2-}{}_{(aq)} \qquad H_2O_{(\ )} + HCO_3{}^-{}_{(aq)}$
acid base base acid

Step 5: $H_3O^+{}_{(aq)} + CO_3{}^{2-}{}_{(aq)} \longrightarrow H_2O_{(\ell)} + HCO_3{}^-{}_{(aq)}$

Reaction 2:

Step 1: bases are $HCO_3{}^-{}_{(aq)}$ and $H_2O_{(\ell)}$.
acids are $H_3O^+{}_{(aq)}$, $HCO_3{}^-{}_{(aq)}$ and $H_2O_{(\ell)}$

Step 2: strongest base is $HCO_3{}^-{}_{(aq)}$

Step 3: strongest acid is $H_3O^+{}_{(aq)}$

Step 4: $H_3O^+{}_{(aq)} + HCO_3{}^-{}_{(aq)} \qquad H_2O_{(\ell)} + H_2CO_{3(aq)}$
acid base base acid

Step 5: $H_3O^+{}_{(aq)} + HCO_3{}^-{}_{(aq)} \longrightarrow H_2O_{(\ell)} + H_2CO_{3(aq)}$

The overall reaction can be obtained by adding the fifth steps of each of Reactions 1 and 2 to obtain

$$2\ H_3O^+{}_{(aq)} + CO_3{}^{2-}{}_{(aq)} \longrightarrow 2\ H_2O_{(\ell)} + H_2CO_{3(aq)}$$

POLYPROTIC ACIDS AND POLYBASIC SPECIES

It should be noted that in some cases the overall reaction cannot be easily obtained since the latter reations in the series may not always go to completion.

Example 1:
A solution of sodium hydroxide is added to phosphoric acid.

Reaction 1:

Step 1: acids are $H_3PO_{4(aq)}$ and $H_2O_{(\ell)}$

bases are $OH^-_{(aq)}$ and $H_2O_{(\ell)}$

Step 2: strongest base is $OH^-_{(aq)}$.

Step 3: strongest acid is $H_3PO_{4(aq)}$.

Step 4: $\underset{\text{acid}}{H_3PO_{4(aq)}} + \underset{\text{base}}{OH^-_{(aq)}} \qquad \underset{\text{acid}}{H_2O_{(\ell)}} + \underset{\text{base}}{H_2PO_4^-{}_{(aq)}}$.

Step 5: since this is a reaction of $OH^-_{(aq)}$ with an acid above HSO_3^- on the ALCHEM data sheet, the reaction is considered quantitative.

$$H_3PO_{4(aq)} + OH^-_{(aq)} \longrightarrow HOH_{(\ell)} + H_2PO_4^-{}_{(aq)}$$

Reaction 2:

Step 1: bases are $OH^-_{(aq)}$, $H_2PO_4^-{}_{(aq)}$ and $H_2O_{(\ell)}$

acids are $H_2PO_4^-{}_{(aq)}$ and $H_2O_{(\ell)}$

Step 2: strongest base is $OH^-_{(aq)}$

Step 3: strongest acid is $H_2PO_{4(aq)}$

Step 4: $\underset{\text{acid}}{H_2PO_4^-{}_{(aq)}} + \underset{\text{base}}{OH^-_{(aq)}} \qquad \underset{\text{base}}{HPO_4^{2-}{}_{(aq)}} + \underset{\text{acid}}{H_2O_{(\ell)}}$

Step 5: The reaction is considered quantitative – $OH^-_{(aq)}$ with acid above $HSO_3^-{}_{(aq)}$.

$$H_2PO_4^-{}_{(aq)} + OH^-_{(aq)} \longrightarrow HPO_4^{2-}{}_{(aq)} + H_2O_{(\ell)}$$

Reaction 3:

Step 1: bases are $OH^-_{(aq)}$, $H_2O_{(\ell)}$, and $HPO_4^{2-}{}_{(aq)}$

acids are $HPO_4^{2-}{}_{(aq)}$ and $H_2O_{(\ell)}$

Step 2: strongest base is $OH^-_{(aq)}$

Step 3: strongest acid is $HPO_4^{2-}{}_{(aq)}$

Step 4: $\underset{\text{base}}{OH^-_{(aq)}} + \underset{\text{acid}}{HPO_4^{2-}{}_{(aq)}} \qquad \underset{\text{acid}}{H_2O_{(\ell)}} + \underset{\text{base}}{PO_4^{3-}{}_{(aq)}}$

Step 5: This is no longer a quantitative reaction ($OH^-_{(aq)}$ with an acid above HSO_3^-).
Since the reactant base is stronger than the product base, equilibrium favors prod

$$OH^-_{(aq)} + HPO_4^{2-}{}_{(aq)} \qquad H_2O_{(\ell)} + PO_4^{3-}{}_{(aq)}$$

It is not proper to write a net equation for all three reactions together since the first two are quantitative and the last is an equilibrium.

POLYPROTIC ACIDS AND POLYBASIC SPECIES

Example 2:

Sodium oxalate is reacted with hydriodic acid ($Na_2OOCCOO$ and $HI_{(aq)}$).

Reaction 1:

Step 1: bases are $OOCCOO^{2-}{}_{(aq)}$ and $H_2O_{(\ell)}$

acids are $H_3O^+{}_{(aq)}$ and $H_2O_{(\ell)}$

spectator ions are $Na^+{}_{(aq)}$ and $I^-{}_{(aq)}$

Step 2: strongest base is $OOCCOO^{2-}{}_{(aq)}$

Step 3: strongest acid is $H_3O^+{}_{(aq)}$

Step 4: $H_3O^+{}_{(aq)} \;+\; OOCCOO^{2-}{}_{(aq)} \qquad\qquad H_2O_{(\ell)} \;+\; HOOCCOO^-{}_{(aq)}$

Step 5: This is a reaction of $H_3O^+{}_{(aq)}$ with a base above NO_2^- on the *ALCHEM* table and therefore quantitative.

$$H_3O^+{}_{(aq)} + OOCCOO^{2-}{}_{(aq)} \longrightarrow H_2O_{(\ell)} + HOOCCOO^-{}_{(aq)}$$

Reaction 2:

Step 1: bases are $HOOCCOO^-{}_{(aq)}$ and $H_2O_{(\ell)}$

acids are $HOOCCOO^-{}_{(aq)}$, $H_3O^+{}_{(aq)}$ and $H_2O_{(\ell)}$

Step 2: strongest base is $HOOCCOO^-{}_{(aq)}$.

Step 3: strongest acid is $H_3O^+{}_{(aq)}$

Step 4: $\underset{\text{base}}{HOOCCOO^-{}_{(aq)}} \;+\; \underset{\text{acid}}{H_3O^+{}_{(aq)}} \qquad\qquad \underset{\text{acid}}{HOOCCOOH_{(aq)}} \;+\; \underset{\text{base}}{H_2O_{(\ell)}}$

Step 5: Reaction 2 does not fall into the category of reaction of $H_3O^+{}_{(aq)}$ with a base bel $NO_2^-{}_{(aq)}$ and therefore is not quantitative. Since the reactant base is stronger than the product base, the equilibrium favored the forward reaction.

$$HOOCCOO^-{}_{(aq)} + H_3O^+{}_{(aq)} \rightleftharpoons H_2O_{(\ell)} \text{ } HOOCCOOH_{(aq)} + H_2O_{(\ell)}$$

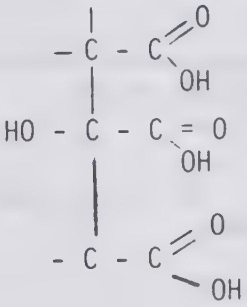
and the two reactions cannot properly be added, since the first reaction is quantitative and the second is not.

ACIDS AND BASES

POLYPROTIC ACIDS AND POLYBASIC SPECIES

For each question, write all reaction steps and where possible the total ionic equatio

1. Hydriodic acid is continuously added to potassium phosphate solution until no more reaction occurs.

2. Citric acid  in lemons is reacted with lithium hydroxide. (all reactions are quantitative)

3. An oxalic acid solution reacts with a barium hydroxide solution.

4. Potassium hydroxide solution is continuously added to sulfuric acid until no more reaction occurs.

STOICHIOMETRIC ACID–BASE TITRATIONS

Acid-base titrations are usually performed to determine the concentration of an acid or base solution. This is most easily done by titrating a sample with a standardized stock laboratory acid or base. The volume of stock solution required to react completely is recorded and a solution stoichiometry calculation is performed as outlined below (same as in Chemistry 20 and in Electrochemistry). However there are differences between redox and acid-base reactions. As outlined in the Unit M, *Electrochemistry*, all redox reactions are stoicheiometric (quantitative) but the problem is that some redox reactions are very slow. Conversely, acid-base reactions are not all stoichiometric (quantitative) but are relatively fast.

It is essential that the reaction performed is quantitative because every stoichiometric calculation assumes complete conversion of reactants to products. Since the most convenient reagents used for titrations are solutions of strong acids and hydroxide solutions (both are inexpensive and common), the rules that were outlined earlier in this unit should be considered before any titration is performed.

1. The reaction of $H_3O^+{}_{(aq)}$ (strong acid) with any base below $NO_2^-{}_{(aq)}$ on the *ALCHEM* acid-base table will be quantitative.
2. the reaction of $OH^-{}_{(aq)}$ with any acid above $HSO_3^-{}_{(aq)}$ on the *ALCHEM* acid–base table will be quantitative.

These rules were formulated from observation of many experimental situations and will successfully predict a quantitative reaction in almost every situation.

Example:

A cleaning solution contains dissolved ammonia, $NH_{3(aq)}$. A 25.0 mℓ sample of the solution is titrated to the endpoint with 0.100 M HCℓ, giving this data:

initial buret reading - 0.6 mℓ
final buret reading - 16.3 mℓ

Step 1: Write a balanced equation for the reaction, using the Bronsted–Lowry steps previously employed.

$$H_3O^+{}_{(aq)} + NH_{3(aq)} \longrightarrow H_2O_{(\ell)} + NH_4^+{}_{(aq)}$$

Step 2: Calculate the # moles of known reagent.

$$\begin{aligned}\#\text{ moles } H_3O^+ &= M \times \#\text{ litres}\\ &= 0.100\ \frac{\text{mol}}{\ell} \times 0.0157\ \ell\\ &= 1.57 \times 10^{-3}\text{ mol}\end{aligned}$$

Step 3: Use the mole ratio from the equation to determine # moles of the reagent of unknown concentration.

$$\begin{aligned}\#\text{ moles } NH_3 &= \#\text{ moles } H_3O^+ \times \text{mol ratio } \frac{NH_3}{H_3O^+}\\ &= 1.57 \times 10^{-3}\text{ mol} \times \frac{1}{1}\\ &= 1.57 \times 10^{-3}\text{ mol}\end{aligned}$$

Step 4: Solve for required information.

$$\begin{aligned}M_{NH_3} &= \frac{\#\text{ moles}}{\#\text{ litres}}\\ &= \frac{1.57 \times 10^{-3}\text{ mol}}{0.0250\ \ell} = 0.0628\text{ M } \textit{or } 6.28 \times 10^{-2}\text{ M}\end{aligned}$$

STOICHIOMETRIC ACID–BASE TITRATIONS

Show all work when solving the following problems.

1. 10.0 ml of 0.400 M KOH solution are required to completely react with 40.0 ml of gastric juice. Find the concentration of the gastric juice ($HC\ell_{(aq)}$).

2. Find the pH of the gastric juice in Question 1.

3. An *ALCHEM* student suspected the vinegar in the cafeteria was diluted. A 10.0 ml sample of the suspected vinegar was titrated to an endpoint with 1.00×10^{-2} M $Ca(OH)_2$. The following data were recorded:

   initial buret reading = 0.4 ml
   final buret reading = 25.4 ml

   Find the concentration of ethanoic acid in the vinegar. Was the vinegar diluted?

STOICHIOMETRIC ACID—BASE TRITRATIONS

4. A sample containing 81.0 g of Sani—flush ($NaHSO_4$) was dissolved in 4.00 ℓ of water in a toilet bowl. The bowl was then titrated with 2.00 M KOH. (It was a titrate-a-bowl sample.) What volume of KOH was required for complete reaction. (The reaction container was later thoroughly flushed with water.)

5. An effluent from a research station was suspected as being phosphoric acid. To determine the concentration of $H_3PO_{4(aq)}$, the solution was titrated with 1.00 M NaOH. Write the equations for the three steps in the reaction.

6. If the size of the sample in Question 5 was 10.0 ml, and 17.3 ml of NaOH were required to completely remove the first proton, find the $[H_3PO_{4(aq)}]$.

7. If a 25.0 ml sample of wine reacted exactly with 45.0 ml of an NaOH solution of pH = 12.0, find the pH of the wine.

TITRATION CURVES – DEMO N4

Purpose:

To plot titration curves for acid-base reactions.

Predemo Information:

Titrations involving acid-base reactions find very wide use in chemistry for analysis of substances of unknown concentrations and for controlling other reactions that depend on pH level.

Definitions

1. *Titration* is the progressive addition of one reagent (usually a liquid) to another.
2. *Stoichiometric acid-base titration* is the addition of one substance to the other until an exact reaction occurs between the acid and the base. In order to stoichiometrically titrate an acid and a base, the reaction must be quantitative. (Recall the rules outlined earlier.) Thus $HC\ell_{(aq)}$ and $NH_{3(aq)}$ can be titrated stoichiometrically but $CH_3COOH_{(aq)}$ and $NH_{3(aq)}$ can not.
3. *Endpoint* in a titration is that point where chemically equivalent reacting amounts have been brought together.
4. *Indicators* are substances that indicate when the end point has been reached. Indicators are selected by choosing one that changes color within the pH range of the end point of the titration (see titration curves).

Although the order is theoretically unimportant a base is usually added to an acid in a titration. A *titration curve* is a graph that indicates change that occurs in pH when a base solution is added to an acid solution. Titration curves are particularly useful for finding the approximate pH at the end point of a titration and for emphasizing that near the endpoint, the change in pH is quite pronounced. The pH at the endpoint will determine which indicator could be chosen if a titration without a pH meter was done.

Materials:

Part A:	Part B:	Part C:
- pH meter	- pH meter	- pH meter
- 50 ml graduated cylinder	- 50 ml graduated cylinder	- 50 ml graduated cylinder
- stock bottle of 0.100 M $HC\ell$ solution	- stock bottle of 0.100 M H_2SO_4 solution	- stock bottle of 0.100 M H_3PO_4 solution
- stock bottle of 0.100 M NaOH solution	- stock bottle of 0.100 M NaOH solution	- stock bottle of 0.100 M NaOH solution
- 150 ml beaker	- 250 ml beaker	- 400 ml beaker
- 10 ml graduated cylinder or pipet	- 10 ml graduated cylinder or pipet	- 10 ml graduated cylinder or pipet

Procedure:

Part A:

1. Calibrate the pH meter with a solution of known $[H_3O^+_{(aq)}]$.
2. Using a graduated cylinder, pour 50.0 ml of 0.100 M $HC\ell$ into a 150 ml beaker.
3. Record the pH of the $HC\ell$ solution in the beaker.
4. Using a graduated cylinder or pipet, add 10.0 ml of 0.100 M NaOH solution to the acid in the beaker.
5. Measure the pH of the mixture.
6. Repeat Steps 4 and 5 until 80 ml of NaOH solution have been added.

ACIDS AND BASES

TITRATION CURVES –DEMO N4

Question:

Part A:

Draw a graph of the titration, plotting the number of ml of NaOH added along the horizontal axis and the pH along the vertical axis.

Procedure:

Part B:

1. Using a graduated cylinder, pour 50.0 ml of 0.100 M $H_2SO_{4(aq)}$ into a 150 ml beaker.
2. Record the pH of the H_2SO_4 solution in the beaker.
3. Using a graduated cylinder or pipet, add 10.0 ml of 0.100 M NaOH solution to the acid in the beaker.
4. Record the pH of the mixture.
5. Repeat Steps 3 and 4 until 120 ml of NaOH solution have been added.

Question:

Part B: Draw a titration curve for the reaction in Part B.

TITRATION CURVES – DEMO N4

Procedure:

Part C:

1. Using a graduated cylinder, pour 50.0 ml of 0.100 M $H_3PO_{4(aq)}$ into a 400 ml beaker.
2. Record the pH of the H_3PO_4 solution in the beaker.
3. Using a graduated cylinder or pipet, add 10.0 ml of 0.100 M NaOH solution to the acid in the beaker.
4. Record the pH of the mixture.
5. Repeat Steps 3 and 4 until 180 ml of NaOH solution have been added.

Question:

Part C:

Draw a titration curve for the reaction in Part C.

STOICHIOMETRIC ACID—BASE TITRATIONS

During a titration, the pH change is small until very near the end-point, where it changes rapidly through a large range. Any indicator that changes color from pH 3 - 11 is suitable for showing the end-point in the titration of $HC\ell_{(aq)}$ and NaOH.

Titrations involving an acid-base reaction have several advantages over other types of stoichiometric analysis. Acid-base titrations:

1. use common reagents
2. use inexpensive labware
3. are easily designed to allow multiple trials
4. are *highly* accurate

Near the end-point for an acid-base reaction, one drop of acid or base in excess can change the pH value by 2 to 3, which is a change in $[H_3O^+_{(aq)}]$ of over 100 to 1000 times. This means that a good indicator will allow measurement of reacting amounts to within part of one drop of solution.

Referring to the previous demo and graph, suppose one drop (about 0.0500 mℓ) of 0.100 M NaOH is added to the reaction after the end-point. Since this reaction is complete (i.e., the same number of moles of $OH^-_{(aq)}$ and $H_3O^+_{(aq)}$ exist), at the end-point the solution is essentially a water solution of $Na^+_{(aq)}$ and $C\ell^-_{(aq)}$ ions, with a pH of 7.00. The one drop of excess NaOH adds 5.00×10^{-6} mol of $OH^-_{(aq)}$.

$$\#\text{ moles} = M \times \#\text{ litres} = 0.100\ \frac{\text{mol}}{\ell} \times 5.00 \times 10^{-5}\ell$$

$$= 5.00 \times 10^{-6}\text{ mol}$$

The excess $OH^-_{(aq)}$ ions remain since there is essentially no $H_3O^+_{(aq)}$ left to react. The pH of the solution may be calculated as follows:

Step 1:

$$[OH^-_{(aq)}] = \frac{\#\text{ moles}}{\#\text{ litres}} = \frac{5.00 \times 10^{-6}\text{ mol}}{0.10005\ \ell} = 5.00 \times 10^{-5}\ M$$

Step 2:

$$[H_3O^+_{(aq)}] = \frac{1.00 \times 10^{-14}}{[OH^-_{(aq)}]} = \frac{1.00 \times 10^{-14}}{5.00 \times 10^{-5}} = 2.00 \times 10^{-10}\ M$$

Step 3:

$$pH = -(\log[H_3O^+_{(aq)}]) = -(\log(2.00 \times 10^{-10}))$$

$$= -(0.30 - 10) = 9.70$$

The one drop of $NaOH_{(aq)}$ increases the volume of OH^- added by only $\frac{1}{1000}$ of the total but the pH increases by 2.70 pH units. This means that $[H_3O^+_{(aq)}]$ has decreased by 500 times, from 1.00×10^{-7} M to 2.00×10^{-10} M. This large change in pH for such a small change in reacting amount gives great sensitivity of measurement.

Most acid-base titrations are done with strong acids and strong bases (like hydroxide ions. According to previous rules establishing quantitative (or stoichiometric) reactions this usually ensures that the reactions are complete and that the pH change at endpoint is large.

STOICHIOMETRIC ACID—BASE TITRATIONS

Polyprotic acids or polybasic species will react in successive steps until all protons have transferred. For example, if sulfuric acid is titrated with potassium hydroxide solution, the reactions are:

Reaction 1:

$$H_3O^+{}_{(aq)} + OH^-{}_{(aq)} \longrightarrow H_2O_{(\ell)} + H_2O_{(\ell)}$$

Possible bases were $OH^-{}_{(aq)}$, $HSO_4^-{}_{(aq)}$ or $H_2O_{(\ell)}$

Possible acids were $H_3O^+{}_{(aq)}$, $HSO_4{}_{(aq)}$ or $H_2O_{(\ell)}$

Reaction 2:

$$HSO_4^-{}_{(aq)} + OH^-{}_{(aq)} \longrightarrow SO_4^{2-}{}_{(aq)} + H_2O_{(\ell)}$$

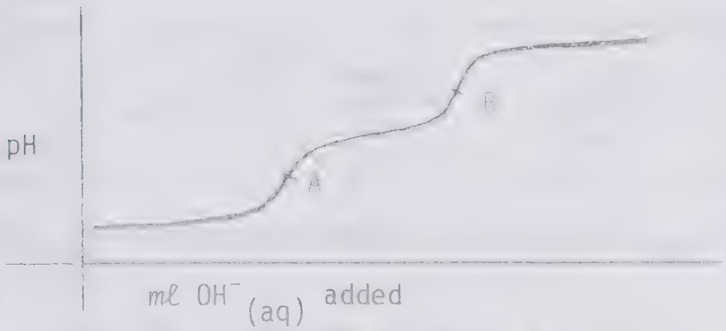
Possible bases were $OH^-{}_{(aq)}$, $HSO_4^-{}_{(aq)}$ or $H_2O_{(\ell)}$

Possible acids were $HSO_4^-{}_{(aq)}$ or $H_2O_{(\ell)}$

The overall reaction is written by combining Reaction 1 and Reaction 2 to give

$$H_3O^+{}_{(aq)} + HSO_4^-{}_{(aq)} + 2\ OH^-{}_{(aq)} \longrightarrow SO_4^{2-}{}_{(aq)} + 3\ H_2O_{(\ell)}$$

The titration curve for the net reaction looks as follows:



Point A is the endpoint for the first step and point B is the endpoint for the entire reaction. Since both endpoints are quantitative, either end-point can be detected with a properly selected indicator or with a pH meter. The latter end-point is usually used in a titration because greater volumes of acid or base added give greater accuracy of measurement.

INDICATORS AND pH DETERMINATION

Indicators

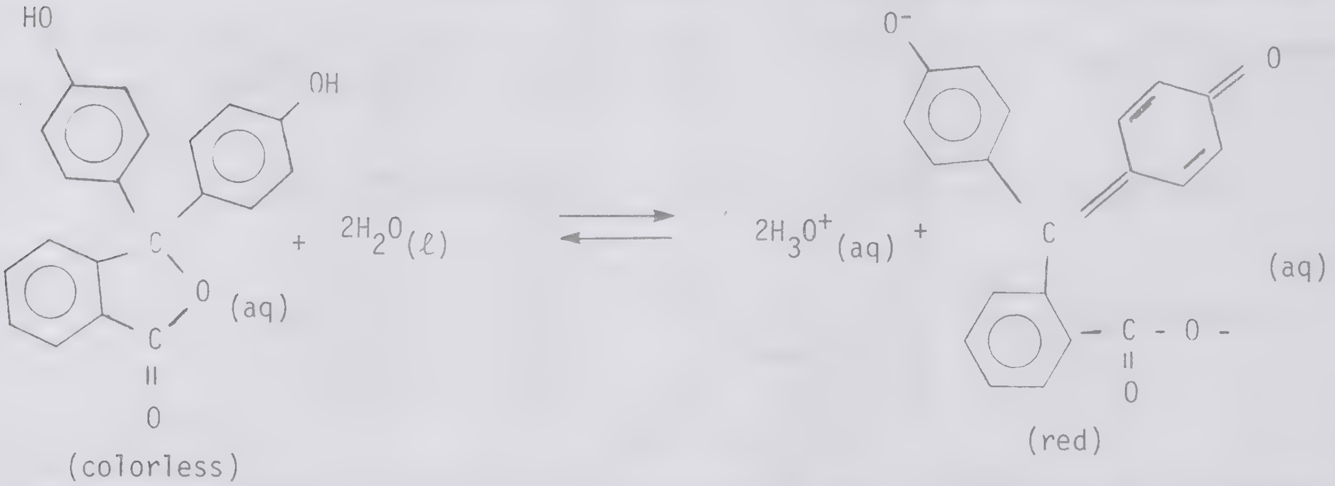
Indicators are substances, usually organic dyes, that react with $H_3O^+_{(aq)}$ or $OH^-_{(aq)}$ in solution to change from one species to another. Since the species have different visible colors, such a substance in solution *indicates* that it is reacting. More importantly, the color change indicates the pH of the solution, since whether or not the reaction occurs depends on pH. The *ALCHEM* data sheet lists some common indicators and their characteristics. The indicator table on the *ALCHEM* data sheet shows that the color change is:

1. not necessarily at the neutral point
2. not at the same pH for any two different indicators
3. not an instant occurrence as the solution pH reaches a specific value.

Indicators are weak acids. Some indicators are included in the *ALCHEM* strength of acids table. Reaction of an indicator can be represented by

$$HIn_{(aq)} + H_2O_{(\ell)} \rightleftharpoons H_3O^+_{(aq)} + In^-_{(aq)}$$

where the acid form HIn and the base form In^- are different colors in solution. (Actually, the structure represented by In^- is usually very complex as in the reaction of phenolphthalein shown below.)



Indicators and Le Chatelier's Principle

When an indicator is added to a strong acid solution, the high $[H_3O^+_{(aq)}]$ due to the strong acid ensures that the equilibrium will react to the left and almost all of the indicator will exist as $HIn_{(aq)}$ (acid form). Conversely when an indicator is added to a strong base solution. the low $[H_3O^+_{(aq)}]$ will ensure that the equilibrium will shift to the right and almost all of the indicator will exist as the $In^-_{(aq)}$ species (base form).

To understand the action of an indicator in a titration, consider adding an indicator to a solution that has a high $[H_3O^+_{(aq)}]$. According to the Le Chatelier Principle, the indicator equilibrium will shift to the left and virtually all of the indicator will gain protons and exist as the $HIn_{(aq)}$ species. If a strong base such as $OH^-_{(aq)}$ is now progressively added to the acid solution, a series of reactions occurs, as predicted by Bronsted-Lowry Theory.

INDICATORS AND pH DETERMINATION

Indicators and Bronsted-Lowry Theory in a Titration

Reaction 1: (the reaction of $OH^-_{(aq)}$ being added to $H_3O^+_{(aq)}$ with an indicator present)

Step 1: bases are $OH^-_{(aq)}$, $H_2O_{(\ell)}$ and $In^-_{(aq)}$
acids are $HIn_{(aq)}$, $H_2O_{(\ell)}$ and $H_3O^+_{(aq)}$

Step 2: strongest base is $OH^-_{(aq)}$

Step 3: strongest acid is $H_3O^+_{(aq)}$

Step 4: $\underset{\text{base}}{OH^-_{(aq)}} + \underset{\text{acid}}{H_3O^+_{(aq)}} \qquad \underset{\text{acid}}{H_2O_{(\ell)}} + \underset{\text{base}}{H_2O_{(\ell)}}$

Step 5: $OH^-_{(aq)} + H_3O^+_{(aq)} \longrightarrow H_2O_{(\ell)} + H_2O_{(\ell)}$

Reaction 2: (the subsequent reaction of $OH^-_{(aq)}$ with the indicator)

Step 1: bases are $OH^-_{(aq)}$, $H_2O_{(\ell)}$ and $In^-_{(aq)}$

acids are $HIn_{(aq)}$ and $H_2O_{(\ell)}$

Step 2: strongest base is $OH^-_{(aq)}$

Step 3: The strongest acid is always $HIn_{(aq)}$. (All common indicators are stronger acids than $H_2O_{(\ell)}$.)

Step 4: $\underset{\text{acid}}{HIn_{(aq)}} + \underset{\text{base}}{OH^-_{(aq)}} \qquad \underset{\text{acid}}{H_2O_{(\ell)}} + \underset{\text{base}}{In^-_{(aq)}}$

Step 5: In most cases, $OH^-_{(aq)}$ is a much stronger base than $In^-_{(aq)}$ and the reaction is essentially complete.

$$HIn_{(aq)} + OH^-_{(aq)} \longrightarrow H_2O_{(\ell)} + In^-_{(aq)}$$

The pH of the color change point thus depends on the strength of In^- *as a proton attractor.* The color change is gradual as $[HIn_{(aq)}]$ decreases and $[In^-_{(aq)}]$ increases. In fact, if both species are colored there will be an intermediate color that appears when the concentration of the two species is roughly equal. For instance, bromthymol blue looks green at this point, because it is about half yellow (acid) species and half blue (base) species.

pH Determination Using Indicators

The pH of an unknown solution may be determined by testing with a series of indicators. For example, if a solution tested blue with bromthymol blue, and colorless with phenolphthalein, its pH would have to be somewhere between about 7.5 - 8.5. Tests of this nature are usually done with pH papers, which are prepared by mixing many indicators, soaking paper in the solution, and letting the paper dry. A strip of such paper will change through a wide range of colors as pH changes. The paper is dipped into the solution to be tested, and the resulting color of paper is compared to a chart. More accurate determinations of pH are usually done with a pH meter, which measures the $[H_3O^+_{(aq)}]$ electrochemically. The $[H_3O^+]$ affects the voltage of the miniature electrochemical cell within the pH meter probe.

STANDARDIZATION TITRATIONS — LAB N2

Purpose:

To accurately determine the concentration of a solution of NaOH by first titrating a primary standard solution of Na_2CO_3 with $HC\ell_{(aq)}$ and then titrating the NaOH solution with the standardized $HC\ell_{(aq)}$.

Prelab Information:

The reagents most often used as acid and base of standard concentration in laboratories are $HC\ell_{(aq)}$ and $NaOH_{(aq)}$. They are both quite stable, inexpensive, and easy to prepare. However, neither is a *primary* standard (i.e., one from which a solution of known concentration can be prepared to high accuracy). $HC\ell$ is gaseous in a pure state and commercially sold solutions vary in concentration. NaOH is a *deliquescent* solid which means it will attract and absorb moisture from the air until it becomes a solution. The absorption of water by NaOH is so rapid that the mass of a sample of NaOH cannot be determined accurately.

In this lab the concentration of solutions of $HC\ell_{(aq)}$ and $NaOH_{(aq)}$ will be determined accurately by first titrating a primary standard solution of Na_2CO_3 with the $HC\ell_{(aq)}$ and then titrating the $NaOH_{(aq)}$ with the standardized $HC\ell_{(aq)}$.

Na_2CO_3 is a *primary* standard because the solid can be obtained in a very pure form and overnight heating will expel any absorbed water vapor. The pure solid Na_2CO_3 can then be used to prepare a solution of known concentration since the mass of a sample of the solid can be measured very accurately.

Materials:

Part A:

- stock supply of $HC\ell_{(aq)}$
- stock bottles of $Na_2CO_{3(s)}$
- 50 ml buret
- wash bottle of distilled water
- buret brush
- liquid soap
- buret stand
- buret clamp
- small funnel
- centigram balance
- 400 ml waste beaker
- 2 - 100 ml beaker
- 2 - 250 ml beaker
- 100 ml volumetric flask with stopper
- 10 ml pipet with bulb
- 250 ml Erlenmeyer flask
- methyl orange indicator
- buret reader
- stirring rod

Part B:

- stock bottles of $HC\ell_{(aq)}$
- stock bottles of $NaOH_{(aq)}$
- 50 ml buret
- buret brush
- liquid soap
- buret stand
- buret clamp
- small funnel
- 100 ml beaker
- 250 ml beaker
- 400 ml waste beaker
- 10 ml pipet with bulb
- bromthymol blue indicator
- buret reader
- 250 ml Erlenmeyer flask

ACIDS AND BASES

STANDARDIZATION TITRATIONS – LAB N2

Procedure:

Part A:

1. Obtain about 100 ml of $HC\ell_{(aq)}$ in a clean, dry, labelled 250 ml beaker.
2. Obtain enough dry, solid Na_2CO_3 to just cover the bottom of a 100 ml beaker.
3. Scrub the buret using the long buret brush. Rinse the buret several times with tap water and at least once with distilled water.
   Note: Handle the buret carefully as the tip is delicate and will break very easily. Do not hold the tip end inside the sink while cleaning.
4. Set up the buret on its stand and set the funnel in the buret. Rinse through into a waste beaker with several 5 ml portions of $HC\ell_{(aq)}$. Close the stopcock and fill the buret with $HC\ell_{(aq)}$ using a funnel. Eliminate any air bubbles from the tip by allowing some of the $HC\ell_{(aq)}$ to run through into the 400 ml waste beaker
5. Find the mass of the second clean, dry 100 ml beaker to the nearest 0.01 g.
6. Set the balance to read exactly 1.50 g more than the mass of the beaker and add $Na_2CO_{3(s)}$ until the beam balances exactly.
7. Add about 50 ml of distilled water to the beaker and dissolve the Na_2CO_3. Using a funnel, carefully transfer the solution into a 100 ml volumetric flask. Rinse the beaker twice using approximately 10 ml of distilled water each time. Pour the rinsings into the volumetric flask.
8. Add distilled water to the 100 ml mark of the volumetric flask, stopper the flask and mix thoroughly. Pour the solution into a clean, dry, labelled 250 ml beaker.
9. Rinse the pipet with a small portion of the Na_2CO_3 solution, then pipet a 10.0 ml sample of this solution into a clean 250 ml Erlenmeyer flask. Add several drops of methyl orange indicator.
10. Record the initial buret reading and titrate with the $HC\ell_{(aq)}$ until the first appearance of a red color. Swirl the flask vigorously to expel any $CO_{2(g)}$ and titrate until the first permanent appearance of a red color.
11. Record the final buret reading.
12. Repeat Steps 9 - 13 at least two more times or until results from three trials are obtained that agree to plus or minus 0.1 ml.
13. If Part B is not to be done in the same lab period, clean the work area. Rinse out all glassware and leave the buret *inverted* with the stopcock *open*.

STANDARDIZATION TITRATION — LAB N2

Observations:

Part A:

Volumes of $HC\ell_{(aq)}$ Used to Neutralize 10.0 ml of $Na_2CO_{3(aq)}$.

Trial	1	2	3	4
Final Buret Reading				
Initial Buret Reading				
Volume of $HC\ell_{(aq)}$ Used				
Average volume of $HC\ell_{(aq)}$ Used =				

Procedure:

Part B:

1. Obtain about 100 ml of $HC\ell_{(aq)}$ in a clean, dry, labelled 250 ml beaker.
2. Obtain about 75 ml of $NaOH_{(aq)}$ in a clean, dry, labelled 100 ml beaker.
3. Clean, set up and fill the buret with $HC\ell_{(aq)}$ as described in Step 4, Part A of the lab.
4. Rinse the pipet with a small portion of the NaOH solution. Pipet a 10.0 ml sample of this solution into a clean 250 ml Erlenmeyer flask. Add several drops of bromthymol blue indicator.
5. Record the initial buret reading and titrate with the $HC\ell_{(aq)}$ until a permanent green color is reached.
   Note: It is extremely easy to overshoot the green color. Try to come as close as possible.
6. Record the final buret reading.
7. Repeat Steps 4 to 6 at least two more times or until results from three trials are obtained that agree to plus or minus 0.1 ml.
8. Clean the work area. Rinse out all glassware and leave the buret inverted with the stopcock open.

Observations:

Part B: Volumes of $HC\ell_{(aq)}$ Used to Neutralize 10.0 ml of $NaOH_{(aq)}$.

Trial #	1	2	3	4
Final Buret Reading				
Initial Buret Reading				
Volume of $HC\ell_{(aq)}$ Used				
Average Volume of $HC\ell_{(aq)}$ Used =				

ACIDS AND BASES

STANDARDIZATION TITRATION – LAB N2

Calculations:

1. Calculate the molar concentration of the Na_2CO_3 solution.

2. Calculate the concentration of the hydrochloric acid (using the volume and molarit of Na_2CO_3 solution and the average volume of $HC\ell_{(aq)}$ required for neutralization.) Show all work.

3. Calculate the concentration of the NaOH solution (using the $[HC\ell_{(aq)}]$ from Question 2 and the data for Part B.)

ACIDS AND BASES

STANDARDIZATION TITRATION — LAB N2

Questions:

1. What is the pH of the hydrochloric acid solution?

2. Find the pH of the sodium hydroxide solution.

Consider the reaction between $HC\ell_{(aq)}$ and $Na_2CO_3{}^{2-}{}_{(aq)}$ when answering Questions 3 - 5.

3. Explain why the reaction must occur in more than one step (although the equation can be written as a one step process).

4. Draw a rough sketch of the titration curve.

5. Explain why methyl orange indicator is used.

ACIDS AND BASES

BUFFERS

A buffer is a solution that resists changes in pH when an acid or a base is added. Buffer solutions are usually composed of a weak acid and the salt of a weak acid. To understand the effect of buffers it is necessary to consider Le Chatelier's Principle.

Example:

How does an acetic acid-sodium acetate solution resist change in pH when an $HC\ell_{(aq)}$ or NaOH solution is added?

In the buffer solution the weak acid exists primarily as molecules (reverse reaction is favoured).

$$\overset{\longleftarrow}{CH_3COOH_{(aq)} + H_2O_{(\ell)} \rightleftharpoons CH_3COO^-_{(aq)} + H_3O^+_{(aq)}}$$

The addition of $OH^-_{(aq)}$ to the buffer will result in some $H_3O^+_{(aq)}$ being consumed by the $OH^-_{(aq)}$:

$$H_3O^+_{(aq)} + OH^-_{(aq)} \longrightarrow H_2O_{(\ell)} + H_2O_{(\ell)}.$$

Since this decreased the concentration of a product, according to Le Chatelier's Principle the equilibrium will try to return the concentration to its original value. The system will shift to the right (more acetic acid molecules will react with water) and the $H_3O^+_{(aq)}$ that was consumed by the $OH^-_{(aq)}$ will be at least partially replaced. Since there is usually a fairly high concentration of acetic acid molecules in the buffer a lot of $H_3O^+_{(aq)}$ can be produced in this manner.

The addition of an $HC\ell_{(aq)}$ solution results in a similar type of reaction. $HC\ell_{(aq)}$ is a strong acid and reacts with water completely.

$$HC\ell_{(aq)} + H_2O_{(\ell)} \longrightarrow H_3O^+_{(aq)} + C\ell^-_{(aq)}$$

The addition of $HC\ell_{(aq)}$ can be considered simply as the addition of $H_3O^+_{(aq)}$ since $C\ell^-_{(aq)}$ is neither an acid nor a base. In the buffer solution the sodium acetate completely dissociates.

$$CH_3COONa_{(s)} \longrightarrow CH_3COO^-_{(aq)} + Na^+_{(aq)}$$

When the $H_3O^+_{(aq)}$ is added to the buffer, it reacts with the base $CH_3COO^-_{(aq)}$ which is present in high concentration from the sodium acetate dissociation.

$$H_3O^+_{(aq)} + CH_3COO^-_{(aq)} \longrightarrow CH_3COOH_{(aq)} + H_2O_{(\ell)}$$

The buffer can continue to absorb $H_3O^+_{(aq)}$ without changing pH until all the acetate ion has reacted.

ACIDS AND BASES

BUFFERS

One of the most important body buffers is the carbonic acid - hydrogen carbonate ion system.

1. Write the equation for the reaction that occurs when $H_3O^+_{(aq)}$ is added to the buffer described above (when a large excess of $H_3O^+_{(aq)}$ is absorbed).

2. Write the equation for the reaction that occurs when $OH^-_{(aq)}$ is added to the buffer described above.

3. Another important body buffer is the $H_2PO_4^-{}_{(aq)}$ - $HPO_4^{2-}{}_{(aq)}$ system. Write an equation for the reaction of the buffer when $NH_{3(aq)}$ is added to the system.

4. Sodium acetate and ethanoic (acetic) acid can form an efficient buffer system. Write an equation for the reaction when stomach acid is added to this buffer.

ACIDS AND BASES

AN OVERVIEW

1. What is the $[H_3O^+]$ in a solution of pH = 9.00?

2. Is the solution in Question 1 acidic or basic?

3. What is the $[OH^-]$ of the solution in Question 1?

Write equations for the reaction between each of the following acid-base pairs. Write an arrow above each set of equilibrium arrows to show whether the forward or reverse reaction was favored as equilibrium was approached.

4. $HNO_{2(aq)} + NH_{3(aq)} \overset{\_\_\_}{\rightleftharpoons}$

5. $H_2SO_{3(aq)} + HCO_3^-{}_{(aq)} \overset{\_\_\_}{\rightleftharpoons}$

6. $HSO_4^-{}_{(aq)} + HPO_4^{2-}{}_{(aq)} \overset{\_\_\_}{\rightleftharpoons}$

7. $H_2PO_4^-{}_{(aq)} + S^{2-}{}_{(aq)} \overset{\_\_\_}{\rightleftharpoons}$

8. $H_3O^+{}_{(aq)} + SO_4^{2-}{}_{(aq)} \overset{\_\_\_}{\rightleftharpoons}$

ACIDS AND BASES

AN OVERVIEW

9. Oil and grease can be removed from concrete floors by a solution of NaOH. If 1.00 g of $NaOH_{(s)}$ is dissolved in water to make 250 ml of solution, what are $[H_3O^+]$, $[OH^-]$ and pH of the solution.

A student prepares the following aqueous solutions.

1.0×10^{-2} M $HC\ell$
1.0 M HF
1.0×10^{-1} M $KHSO_4$
1.0 M H_2S

10. Which solution contains the strongest acid?

11. Which solution contains the weakest acid?

12. If all four solutions were equally concentrated, how would they be ranked from highest pH to lowest pH?

13. If all four solutions were equally concentrated how would they be ranked from highest $[H_3O^+]$ to lowest $[H_3O^+]$?

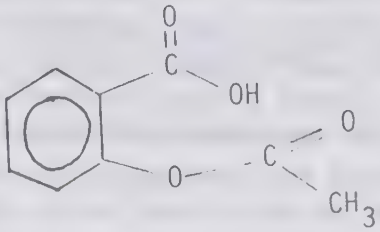
14. Use the % dissociation information from the *ALCHEM* data sheet to calculate $[H_3O^+]$ for each solution assuming each is 1.0 M and rank them from most to least acidic.

ACIDS AND BASES

AN OVERVIEW

Use the following information to answer Questions 15 to 18.

A student dissolves 59.7 g of acetylsalicylic acid (aspirin) in water to make 1.00 ℓ of solution.

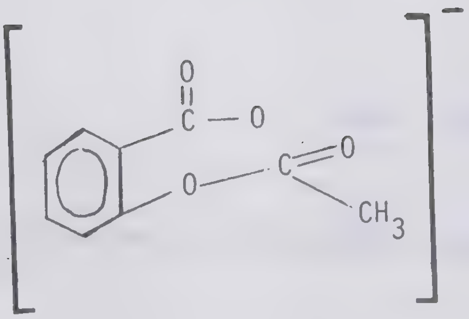


15. Find the concentration of the solution.

16. If the solution is tested and found to have a pH of 2.52 calculate the $[H_3O^+]$ in solution.

17. What is the % dissociation of ASA in the solution?

18. What is



in this solution?

ACIDS AND BASES

AN OVERVIEW

Use the following information to answer Questions 19 to 22.

If a 10.00 g zinc strip is placed in one litre of 0.020 M methanoic acid solution, it will react very slowly, but the reaction will proceed until complete. If a similar Zn strip is placed in one litre of 0.020 M $HC\ell$, the reaction proceeds much more quickly, but exactly the same amount of Zn will be consumed.

19. Explain the differences in the rates of reaction. (What is the species that the Zn atoms actually react with acid molecules or hydronium ions?)

20. What happens to the equilibrium of methanoic acid as the methanoic acid solution reacts with Zn?

21. What mass of Zn is consumed in each ease?

22. What is the final solution composed of in each case?

ACIDS AND BASES

AN OVERVIEW

Use the following information to answer Questions 23 and 24.

A stoichiometric acid-base titration is performed by titrating 100 ml of 0.200 M benzoic acid with a KOH solution of unknown concentration. 25.0 ml of the base solution are required for reaction to endpoint.

23. Write and balance the equation for the reaction and calculate the molarity of the KOH solution.

24. Assuming that the solution contains K^+ and $C_6H_5COO^-$ ions at endpoint, would you expect the pH to be greater or less than 7.0?
(Hint: Are there acids or bases present to increase or decrease $[H_3O^+]$ found in pure water?)

Use the following information to answer Question 25 - 27.

A 500 ml sample of 2.00 M ammonia solution is neutralized by 500 ml of 2.00 M $HCl_{(aq)}$, according to the reaction

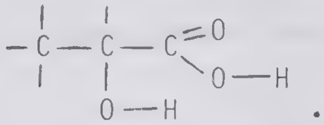
$$H_3O^+{}_{(aq)} + NH_{3(aq)} \longrightarrow NH_4{}^+{}_{(aq)} + H_2O_{(\ell)}$$

25. Assuming complete reaction, what is $[NH_4{}^+]$ in the solution at endpoint?

26. What will be the pH at endpoint assuming that the product $NH^+_{4(aq)}$ reacts with water to the extent of 2.4×10^{-3} %.

27. Which indicator on the *ALCHEM* data sheet would be best for this reaction if done as a titration.

28. Lactic acid builds up in muscle tissue during exercise. The presence of the acid creates the tired, aching feeling that results when a muscle is overexerted. The compound has the structural formula



Write a balanced equation for the reation that occurs when lactic acid comes in contact with a carbonic acid - sodium hydrogen carbonate buffer.

29. Limestone, like marble and chalk, is a crystalline form of calcium carbonate. When these substances come in contact with CO_2 in aqueous solution ($H_2CO_{3(aq)}$) they react to form soluble calcium hydrogen carbonate (an acid-base reaction). Write a balanced equation for the dissolving of limestone in a rainwater solution containing dissolved CO_2 (Ca^{2+} is a spectator ion in the reaction).

30. The calcium hydrogen carbonate solution formed in Question 29 sometimes drips into caves. As the solution evaporates ($CO_{2(g)}$ + $H_2O_{(g)}$ leave) calcium carbonate solidifies in icicles shapes known as stalactites (from the ceiling) and stalagmites (from the floor). Write a balanced equation for the formation of stalagmites.

31. Write the balanced equation for the reaction of a lye solution with an eyewash of $H_3BO_{3(aq)}$. Show each step in the reaction and a net ionic equation.

ACIDS AND BASES

AN OVERVIEW

Many remedies for excess stomach acid have been tried. Write Bronsted-Lo [illegible]
reactions for each remedy in Question 32, 33 and 34.

32. Milk of Magnesia

33. Sodium bicarbonate

Note:

advantages - *cheap, immediate effect in the stomach.*

disadvantages - *makes you belch and if too much is taken, the excess* [illegible] *the intestine where it is absorbed into the bloostrea* [illegible] *interferes with the delicate* H_2CO_3 - HCO_3^- *equilibr* [illegible]

makes the blood more basic (alkalosis).

34. Rolaids. $Al\,Na(OH)_2CO_3$ (Products are aluminum chloride, sodium chloride, [illegible] carbon dioxide).

Note: *Assuming each tablet is pure* $AlNa(OH)_2CO_3$, *calculate the mass* [illegible] *(density = 1.0 g/ml) that will react with 1.0 mol of the soli* [illegible] *ratio of the masses of the two reactants. Can Rolaids consum* [illegible] *weight in excess stomach acid"? (Notice-you are comparing the* [illegible] *with the mass of a dilute solution.*

35. Phosphoric acid is reacted with aqueous ammonia to produce ammonium phospha [illegible]
Write a balanced equation for each step.

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